Chemistry

A sign in a university professor's office asks, What in the world isn't chemistry? Although meant to amuse, this question has a world of truth behind it. High school students come into contact with chemistry every day, often without realizing it. Discussions of daily interactions with chemistry often provide an entry into teaching the subject in high school. Although relating chemistry to daily life is helpful, this approach does not diminish the need for students to have a high level of readiness before entering the class. Of paramount importance is a firm grounding in algebra.

Chemistry is a sequential, hierarchical science that is descriptive and theoretical. It requires knowing the macroscopic properties of matter and the microscopic properties of matter's constituent particles. Although chemical demonstrations may engage students, going beyond a superficial appreciation of chemistry is a critical step. Chemistry requires high-level problem-solving skills, such as designing experiments and solving word problems. For students to learn concepts of chemistry, they must learn new vocabulary, including the rules for naming simple compounds and ions.

Students can discover chemistry's tremendous power to explain the nature of matter and its transformations when they study the periodic table of the elements. Students who move beyond a trivial treatment of the discipline can explore the many useful, elegant, and even beautiful aspects of chemistry. Bringing students to this understanding is a great achievement.

STANDARD SET 1. Atomic and Molecular Structure

Starting with atomic and molecular structure in the study of high school chemistry is important because this topic is a foundation of the discipline. However, because structural concepts are highly theoretical and deal with the quantum realm, they can be hard to relate to real-world experience. Ideally, from grades three through eight, students have been gradually introduced to the atomic theory; and by the end of the eighth grade, they should have covered the major concepts in the structure of atoms and molecules. By the time students reach high school, they should be familiar with basic aspects of this theory.

The study of structure can begin with the simplest element, hydrogen. Students can progress from a simple model of the atom (see standards 1.a, 1.d, 1.e, 1.h*, and 11.a in this section) to the historic Bohr model (see Standard 1.i* in this section) and, finally, to a quantum mechanical model (see standards 1.g* and 1.j* in this section), which is the picture of the atom that students should ultimately develop. Students should learn that the quantum mechanical model takes into account

the particle and wave properties of the electron and uses mathematical equations to solve for electron energies and regions of electron density.

Students should understand that the energy carried by electrons either within an atom or as electricity can be transformed into light energy. Those who have completed high school physics will be familiar with the properties of electromagnetic waves. In chemistry students learn to apply the equations $E = h\nu$ and $c = \lambda\nu$. Students without significant training in physics will need to understand electromagnetic radiation as energy, frequency, and wavelength. The necessary mathematical background for the study of chemistry includes algebraic isolation of variables, use of conversion factors, and manipulation of exponents, all of which are covered in the mathematics standards for the middle grades.

1. The periodic table displays the elements in increasing atomic number and shows how periodicity of the physical and chemical properties of the elements relates to atomic structure. As a basis for understanding this concept:

a. *Students know* how to relate the position of an element in the periodic table to its atomic number and atomic mass.

An atom consists of a nucleus made of protons and neutrons that is orbited by electrons. The number of protons, not electrons or neutrons, determines the unique properties of an element. This number of protons is called the element's atomic number. Elements are arranged on the periodic table in order of increasing atomic number. Historically, elements were ordered by atomic mass, but now scientists know that this order would lead to misplaced elements (e.g., tellurium and iodine) because differences in the number of neutrons for isotopes of the same element affect the atomic mass but do not change the identity of the element.

1. b. *Students know* how to use the periodic table to identify metals, semimetals, nonmetals, and halogens.

Most periodic tables have a heavy stepped line running from boron to astatine. Elements to the immediate right and left of this line, excluding the metal aluminum, are semimetals and have properties that are intermediate between metals and nonmetals. Elements further to the left are metals. Those further to the right are nonmetals. Halogens, which are a well-known family of nonmetals, are found in Group 17 (formerly referred to as Group VIIA). A group, also sometimes called a "family," is found in a vertical column in the periodic table.

1. c. *Students know* how to use the periodic table to identify alkali metals, alkaline earth metals and transition metals, trends in ionization energy, electronegativity, and the relative sizes of ions and atoms.

A few other groups are given family names. These include the alkali metals (Group 1), such as sodium and potassium, which are soft and white and extremely

reactive chemically. Alkaline earth metals (Group 2), such as magnesium and calcium, are found in the second column of the periodic table. The transition metals (Groups 3 through 12) are represented by some of the most common metals, such as iron, copper, gold, mercury, silver, and zinc. All these elements have electrons in their outer *d* orbitals.

Electronegativity is a measure of the ability of an atom of an element to attract electrons toward itself in a chemical bond. The values of electronegativity calculated for various elements range from one or less for the alkali metals to three and one-half for oxygen to about four for fluorine. *Ionization energy* is the energy it takes to remove an electron from an atom. An element often has multiple ionization energies, which correspond to the energy needed to remove first, second, third, and so forth electrons from the atom. Generally in the periodic table, ionization energy and electronegativity increase from left to right because of increasing numbers of protons and decrease from top to bottom owing to an increasing distance between electrons and the nucleus. Atomic and ionic sizes generally decrease from left to right and increase from top to bottom for the same reasons. Exceptions to these general trends in properties occur because of filled and half-filled subshells of electrons.

1. d. *Students know* how to use the periodic table to determine the number of electrons available for bonding.

Only electrons in the outermost energy levels of the atom are available for bonding; this outermost bundle of energy levels is often referred to as the *valence shell* or *valence shell of orbitals.* All the elements in a group have the same number of electrons in their outermost energy level. Therefore, alkali metals (Group 1) have one electron available for bonding, alkaline earth metals (Group 2) have two, and elements in Group 13 (once called Group III) have three. Unfilled energy levels are also available for bonding. For example, Group 16, the chalcogens, has room for two more electrons; and Group 17, the halogens, has room for one more electron to fill its outermost energy level.

To find the number of electrons available for bonding or the number of unfilled electron positions for a given element, students can examine the combining ratios of the element's compounds. For instance, one atom of an element from Group 2 will most often combine with two atoms of an element from Group 17 (e.g., $MgCl_2$) because Group 2 elements have two electrons available for bonding, and Group 17 elements have only one electron position open in the outermost energy level. (Note that some periodic tables indicate an element's electron configuration or preferred oxidation states. This information is useful in determining how many electrons are involved in bonding.)

1. e. *Students know* the nucleus of the atom is much smaller than the atom yet contains most of its mass.

The volume of the hydrogen nucleus is about one trillion times less than the volume of the hydrogen atom, yet the nucleus contains almost all the mass in the

form of one proton. The diameter of an atom of any one of the elements is about 10,000 to 100,000 times greater than the diameter of the nucleus. The mass of the atom is densely packed in the nucleus.

The electrons occupy a large region of space centered around a tiny nucleus, and so it is this region that defines the volume of the atom. If the nucleus (proton) of a hydrogen atom were as large as the width of a human thumb, the electron would be on the average about one kilometer away in a great expanse of empty space. The electron is almost 2,000 times lighter than the proton; therefore, the large region of space occupied by the electron contains less than 0.1 percent of the mass of the atom.

1. f.* *Students know* how to use the periodic table to identify the lanthanide, actinide, and transactinide elements and know that the transuranium elements were synthesized and identified in laboratory experiments through the use of nuclear accelerators.

The lanthanide series, or rare earths, and the actinide series, all of which are radioactive, are separated for reasons of practical display from the main body of the periodic table. If these two series were inserted into the main body, the table would be wider by 14 elements and less manageable for viewing. Within each of these series, properties are similar because the configurations of outer electrons are similar. As a general rule elements in both series appear to have three electrons available for bonding. They combine with halogens to form compounds with the general formula MX_3, such as $LaFz_3$. The transactinide elements begin with rutherfordium, element 104.

All the elements with atomic numbers greater than 92 were first synthesized and identified in experiments. These experiments required the invention and use of accelerators, which are electromagnetic devices designed to create new elements by accelerating and colliding the positively charged nuclei of atoms. Ernest O. Lawrence, at the University of California, Berkeley, invented one of the most useful nuclear accelerators, the cyclotron. Many transuranium elements, such as 97-berkelium, 98-californium, 103-lawrencium, and 106-seaborgium, were first created and identified at the adjacent Lawrence Berkeley National Laboratory. Today a few transuranium elements are produced in nuclear reactors. For example, hundreds of metric tons of plutonium have been produced in commercial nuclear reactors.

1. g.* *Students know* how to relate the position of an element in the periodic table to its quantum electron configuration and to its reactivity with other elements in the table.

Each element has a unique electron configuration (also known as quantum electron configuration) that determines the properties of the element. Quantum mechanical calculations predict these electron energy states, which provide the theoretical justification for the organization of the periodic table, previously organized on the basis of chemical properties.

Students can learn the principal quantum numbers—which are 1, 2, 3, 4, 5, 6, and 7—and the corresponding periods, or horizontal rows, on the periodic table. They can learn the angular momentum quantum numbers that give rise to s, p, d, and f subshells of orbitals and the rules for the sequence of orbital filling. The electrons in the highest energy orbitals with the same principal quantum number are the *valence electrons.* For example, for all elements in Group 1, the valence electron configuration is ns^1, where n is the principal quantum number. Analogously, all elements in Group 16 have valence electron configurations of ns^2np^4. Particular configurations of valence electrons are associated with regular patterns in chemical reactivity. Generally, those elements with one electron in excess or one electron short of a full octet in the highest occupied energy level, the alkali metals and halogens, respectively, are the most reactive.

1. h.* *Students know* the experimental basis for Thomson's discovery of the electron, Rutherford's nuclear atom, Millikan's oil drop experiment, and Einstein's explanation of the photoelectric effect.

In 1887 J. J. Thomson performed experiments from which he concluded that cathode rays are streams of negative, identical particles, which he named *electrons.* In 1913 E. R. Rutherford headed a group that shot a beam of helium nuclei, or alpha particles, through a very thin piece of gold foil; the unexpectedly large deflections of some helium nuclei led to the hypothesis that the charged mass of each gold atom is concentrated in a small central nucleus. Robert Millikan confirmed Thomson's conclusion that electrons are identical particles when he balanced tiny, electrically charged oil droplets between electric and gravitational fields and so discovered that the droplets always contained charge equal to an integral multiple of a single unit.

Albert Einstein found he could explain the photoelectric effect, in which light ejects electrons from metal surfaces, by proposing that light consists of discrete bundles of energy, or *photons,* each photon with an energy directly proportional to the frequency of the light, and by proposing that each photon could eject one and only one electron. Photons with sufficient energy will eject an electron whose kinetic energy is equal to the photon energy minus the energy required to free the electron from the metal. If the frequency of the light and therefore the energy of each proton is too low to free an electron, then merely increasing the light's intensity (that is, merely producing more photons) does not cause electrons to eject.

1. i.* *Students know* the experimental basis for the development of the quantum theory of atomic structure and the historical importance of the Bohr model of the atom.

Niels Bohr combined the concepts of Rutherford's nuclear atom and Einstein's photons with several other ideas to develop a model that successfully explains the observed spectrum, or wavelengths, of electromagnetic radiation that is emitted when a hydrogen atom falls from a high energy state to a low energy state. In classical physics all accelerating charges emit energy. If electrons in an atom behaved in

this way, light of ever-decreasing frequencies would be emitted from atoms. Bohr explained why this phenomenon does not occur when he suggested that electrons in atoms gain or lose energy only by making transitions from discrete energy levels. This idea was a key to the development of nonclassical descriptions of atoms. Louis de Broglie advanced the understanding of the nature of matter by proposing that particles have wave properties. On the basis of these ideas, Erwin Schrödinger and others developed quantum mechanics, a theory that describes and predicts atomic and nuclear phenomena.

1. j.* *Students know* that spectral lines are the result of transitions of electrons between energy levels and that these lines correspond to photons with a frequency related to the energy spacing between levels by using Planck's relationship *(E = hv)*.

The Bohr model gives a simple explanation of the spectrum of the hydrogen atom. An electron that loses energy in going from a higher energy level to a lower one emits a photon of light, with energy equal to the difference between the two energy levels. Transitions of electrons from higher energy states to lower energy states yield *emission,* or *bright-line, spectra. Absorption spectra* occur when electrons jump to higher energy levels as a result of absorbing photons of light. When atoms or molecules absorb or emit light, the absolute value of the energy change is equal to hc/λ, where h is a number called *Planck's constant, c* is the speed of light, and λ is the wavelength of light emitted, yielding Planck's relationship $E = hv$.

The transition from the Bohr model to the quantum mechanical model of the atom requires students to be aware of the probabilistic nature of the distribution of electrons around the atom. A "dart board" made of concentric rings can serve as a two-dimensional model of the three-dimensional atom. A graph as a function of radius of the number of random hits by a dart can be compared with a similar graph as a function of radius of the probability density of the electron in a hydrogen atom.

STANDARD SET 2. Chemical Bonds

Standard Set 2 deals with two distinct topics: chemical bonds and intermolecular attractive forces, such as hydrogen bonds. A logical place to begin the study of this standard is with a discussion of the chemical bond. A key point to emphasize is that when atoms of two different elements join to form a covalent bond, energy is almost always released. Conversely, breaking bonds always requires the addition of energy. Students should understand that the sum of these two processes determines the net energy released or absorbed in a chemical reaction.

This standard set requires a basic knowledge of electrostatics and electronegativity and a thorough knowledge of the periodic table. After studying standards for chemistry for the elementary grades, students should know that matter is made of atoms and that atoms combine to form molecules. Students can also be expected to

know that atoms consist of protons, neutrons, and electrons. Although knowledge of complex mathematics is not required for this standard, some background in three-dimensional geometry will be helpful.

2. Biological, chemical, and physical properties of matter result from the ability of atoms to form bonds from electrostatic forces between electrons and protons and between atoms and molecules. As a basis for understanding this concept:

a. *Students know* atoms combine to form molecules by sharing electrons to form covalent or metallic bonds or by exchanging electrons to form ionic bonds.

In the localized electron model, a covalent bond appears as a shared pair of electrons contained in a region of overlap between two atomic orbitals. Atoms (usually nonmetals) of similar electronegativities can form covalent bonds to become molecules. In a covalent bond, therefore, bonding electron pairs are localized in the region between the bonded atoms. In metals valence electrons are not localized to individual atoms but are free to move to temporarily occupy vacant orbitals on adjacent metal atoms. For this reason metals conduct electricity well.

When an electron from an atom with low electronegativity (e.g., a metal) is removed by another atom with high electronegativity (e.g., a nonmetal), the two atoms become oppositely charged ions that attract each other, resulting in an ionic bond. Chemical bonds between atoms can be almost entirely covalent, almost entirely ionic, or in between these two extremes. The triple bond in nitrogen molecules (N_2) is nearly 100 percent covalent. A salt such as sodium chloride (NaCl) has bonds that are nearly completely ionic. However, the electrons in gaseous hydrogen chloride are shared somewhat unevenly between the two atoms. This kind of bond is called *polar covalent.*

(Note that elements in groups 1, 2, 16, and 17 in the periodic table usually gain or lose electrons through the formation of either ionic or covalent bonds, resulting in eight outer shell electrons. This behavior is sometimes described as "the octet rule.")

2. b. *Students know* chemical bonds between atoms in molecules such as H_2, CH_4, NH_3, H_2CCH_2, N_2, Cl_2, and many large biological molecules are covalent.

Organic and biological molecules consist primarily of carbon, oxygen, hydrogen, and nitrogen. These elements share valence electrons to form bonds so that the outer electron energy levels of each atom are filled and have electron configurations like those of the nearest noble gas element. (Noble gases, or inert gases, are in the last column on the right of the periodic table.) For example, nitrogen has one lone pair and three unpaired electrons and therefore can form covalent bonds with three hydrogen atoms to make four electron pairs around the nitrogen. Carbon has four unpaired electrons and combines with hydrogen, nitrogen, and oxygen to form covalent bonds sharing electron pairs.

The great variety of combinations of carbon, nitrogen, oxygen, and hydrogen make it possible, through covalent bond formation, to have many compounds from just these few elements. Teachers can use ball and stick or gumdrop and toothpick models to explore possible bonding combinations.

2. c. *Students know* salt crystals, such as NaCl, are repeating patterns of positive and negative ions held together by electrostatic attraction.

The energy that holds ionic compounds together, called *lattice energy,* is caused by the electrostatic attraction of *cations,* which are positive ions, with *anions,* which are negative ions. To minimize their energy state, the ions form repeating patterns that reduce the distance between positive and negative ions and maximize the distance between ions of like charges.

2. d. *Students know* the atoms and molecules in liquids move in a random pattern relative to one another because the intermolecular forces are too weak to hold the atoms or molecules in a solid form.

In any substance at any temperature, the forces holding the material together are opposed by the internal energy of particle motion, which tends to break the substance apart. In a solid, internal agitation is insufficient to overcome intermolecular or interatomic forces. When enough energy is added to the solid, the kinetic energy of the atoms and molecules increases sufficiently to overcome the attractive forces between the particles, and they break free of their fixed lattice positions. This change, called *melting,* forms a liquid, which is disordered and nonrigid. The particles in the liquid are free to move about randomly although they remain in contact with each other.

2. e. *Students know* how to draw Lewis dot structures.

A Lewis dot structure shows how valence electrons and covalent bonds are arranged between atoms in a molecule. Teachers should follow the rules for drawing Lewis dot diagrams provided in a chemistry textbook. Students should be able to use the periodic table to determine the number of valence electrons for each element in Groups 1 through 3 and 13 through 18. Carbon, for example, would have four valence electrons. Lewis dot diagrams represent each electron as a dot or an *x* placed around the symbol for carbon, which is C. A covalent bond is shown as a pair of dots, or *x*'s, representing a pair of electrons. For example, a Lewis dot diagram for methane, which is CH_4, would appear as shown in Figure 3.

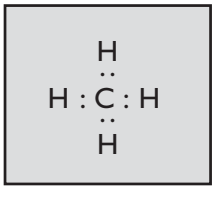


Fig. 3. Lewis Dot Diagram

Lewis dot diagrams provide a method for predicting correct combining ratios between atoms and for determining aspects of chemical bonds, such as whether they are covalent or consist of single, double, or triple bonds.

2. f.* *Students know* how to predict the shape of simple molecules and their polarity from Lewis dot structures.

Using information obtained from Lewis dot structures of covalently bonded molecules, students can predict the overall geometry of those molecules. This model assumes that valence electron pairs repel each other and that atoms covalently bonded around a central atom position themselves as far apart as possible while maintaining the covalent bond. The model also assumes that double or triple bonds define a single electronic region.

To predict shapes, students start with a correct Lewis dot structure. From the number of electron pairs or regions, both bonded and nonbonded, students can determine the molecular geometry of the molecule because Lewis dot structures, although drawn in two dimensions, represent the three-dimensional symmetry of the molecule. A symmetrical distribution of the electron clouds around a central atom leads to a nonpolar molecule in which charge is evenly distributed. Students should be able to predict that a molecule such as methane, with one carbon and four hydrogen atoms, forms a tetrahedral shape and because of its symmetry is a nonpolar molecule.

2. g.* *Students know* how electronegativity and ionization energy relate to bond formation.

During bond formation atoms with large electronegativity values, such as fluorine and oxygen, attract electrons away from lower electronegativity atoms, such as the alkali metals. The difference in electronegativity between the two bonding atoms gives information on how evenly an electron pair is shared. A large difference in electronegativity leads to an ionic bond, with essentially no sharing of electrons. This phenomenon usually occurs between metal and nonmetal atoms. A small difference in electronegativity leads to a covalent bond with more equal sharing of electrons. This result typically occurs between two nonmetal atoms. Electronegativity is related to *ionization energy,* the energy needed to remove an electron from an isolated gaseous atom, leaving a positively charged ion. High ionization energies usually correlate with large electronegativities.

2. h.* *Students know* how to identify solids and liquids held together by van der Waals forces or hydrogen bonding and relate these forces to volatility and boiling/melting point temperatures.

Liquids and solids that are held together not by covalent or ionic bonds but only by weaker intermolecular forces tend to have low to moderate melting and boiling points and to be from somewhat to very volatile. The *volatility* of a substance means how readily it evaporates at ordinary temperatures and pressures.

Two kinds of intermolecular forces are hydrogen bonding and van der Waals attractions (often referred to as London dispersion forces). Hydrogen bonding is essential to life and gives water many of its unusual properties. A hydrogen bond occurs when a hydrogen atom on one molecule, bonded directly to a highly electronegative atom (fluorine, oxygen, or nitrogen), is weakly attracted to the electronegative atom on a neighboring molecule. In the important case of water, this attraction exists between the hydrogen on one water molecule and the oxygen on a neighboring water molecule. This attraction happens because of water's polar nature and bent shape.

Van der Waals forces exist between all molecules, polar or nonpolar. Even when the molecule is nonpolar, the electrons move around and may sometimes find themselves temporarily closer to one nucleus than to the other. The atom with the greater share of the electron density becomes, for an instant, slightly negatively charged, and the other atom becomes a little bit positive. If the same thing happens in a nearby molecule, the positive and negative centers on the two molecules temporarily attract each other. Bigger molecules have more electrons, and their van der Waals forces are greater. This phenomenon leads to molecules with lower volatility and higher melting and boiling temperatures.

STANDARD SET 3. Conservation of Matter and Stoichiometry

Standard Set 3 demands more facility with mathematics than do the previous two chemistry standard sets. For this reason students need prerequisite mathematical skills in two broad categories. First, they must be able to manipulate very large and very small numbers by using exponents as expressed in scientific notation, and they should learn the rules of significant digits when reporting measurements and the results of calculations. Second, students must be able to manipulate simple equations in symbolic form, such as the isolation of variables, and to write equations numerically with correct units. Handling units successfully is necessary for problems involving mole-to-mass and mass-to-mole conversions. The ability to make other types of unit conversion and to square and cube linear measurements will also be required. An understanding of ratios will help students to see the logic behind problems related to Standard Set 3.

Simple metric conversions are relatively easy to grasp, and they contain the basic elements of stoichiometric calculations. For example, to convert 510 nanometers to meters, students write the unknown unit, which in this case is meters; set it equal to the given unit, which in this case is nanometers; and then multiply by the correct conversion factor, making sure that the desired unit is in the numerator and the unit to be cancelled is in the denominator. The following equation demonstrates this procedure:

$$510 \text{ nm} = 510 \text{ nm} \times (10^{-9} \text{ m}/1 \text{ nm}) = 510 \times 10^{-9} \text{ m} = 5.1 \times 10^{-7} \text{ m}$$

This technique converts any units however complicated. For example, centimeters can be converted to nanometers by multiplying by two factors: one that converts from centimeters to meters and another that converts from meters to nanometers. Converting to an area requires the square of a conversion factor; converting to a volume requires the cube. Once students have tackled simple problems, they can move on to more complex stoichiometric relations. They will also need to learn to balance equations easily.

3. The conservation of atoms in chemical reactions leads to the principle of conservation of matter and the ability to calculate the mass of products and reactants. As a basis for understanding this concept:

a. *Students know* how to describe chemical reactions by writing balanced equations.

Reactions are described by balanced equations because all the atoms of the reactants must be accounted for in the reaction products. An equation with all correct chemical formulas can be balanced by a number of methods, the simplest being by inspection. Given an unbalanced equation, students can do an inventory to determine how many of each atom are on each side of the equation. If the result is not equal for all atoms, coefficients (not subscripts) are changed until balance is achieved. Sometimes, reactions refer to substances with written names rather than to chemical symbols. Students should learn the rules of chemical nomenclature. This knowledge can be acquired in stages as new categories of functional groups are introduced.

3. b. *Students know* the quantity *one mole* is set by defining one mole of carbon-12 atoms to have a mass of exactly 12 grams.

The mole concept is often difficult for students to understand at first, but they can be taught that the concept is convenient in chemistry just as a dozen is a convenient concept, or measurement unit, in the grocery store. The mole is a number. Specifically, a *mole* is defined as the number of atoms in 12 grams of carbon-12. When atomic masses were assigned to elements, the mass of 12 grams of carbon-12 was selected as a standard reference to which the masses of all other elements are compared. The number of atoms in 12 grams of carbon-12 is defined as one *mole,* or conversely, if one mole of ^{12}C atoms were weighed, it would weigh exactly 12 grams. (Note that carbon, as found in nature, is a mixture of isotopes, including atoms of carbon-12, carbon-13, and trace amounts of carbon-14.) The definition of the mole refers to pure carbon-12.

The *atomic mass* of an element is the weighted average of the mass of one mole of its atoms based on the abundance of all its naturally occurring isotopes. The atomic mass of carbon is 12.011 grams. If naturally occurring carbon is combined with oxygen to form carbon dioxide, the mass of one mole of naturally occurring oxygen can be determined from the combining mass ratios of the two elements. For example, the weight, or atomic mass, of one mole of oxygen containing mostly oxygen-16 and a small amount of oxygen-18 is 15.999 grams.

3. c. *Students know* one mole equals 6.02×10^{23} particles (atoms or molecules).

A mole is a very large number. Standard 3.b describes the mole as the number of atoms in 12 grams of ^{12}C. The number of atoms in a mole has been found experimentally to be about 6.02×10^{23}. This number, called Avogadro's number, is known to a high degree of accuracy.

3. d. *Students know* how to determine the molar mass of a molecule from its chemical formula and a table of atomic masses and how to convert the mass of a molecular substance to moles, number of particles, or volume of gas at standard temperature and pressure.

The molar mass of a compound, which is also called either the *molecular mass* or *molecular weight,* is the sum of the atomic masses of the constituent atoms of each element in the molecule. Molar mass is expressed in units of grams per mole. The periodic table is a useful reference for finding the atomic masses of each element. For example, one mole of carbon dioxide molecules contains one mole of carbon atoms weighing 12.011 grams and two moles of oxygen atoms weighing 2×15.999 grams for a total molecular mass of 44.009 grams per mole of carbon dioxide molecules.

The mass of a sample of a compound can be converted to moles by dividing its mass by the molar mass of the compound. This process is similar to the unit conversion discussed in the introduction to Standard Set 3. The number of particles in the sample is determined by multiplying the number of moles by Avogadro's number. The volume of an ideal or a nearly ideal gas at a fixed temperature and pressure is proportional to the number of moles. Students should be able to calculate the number of moles of a gas from its volume by using the relationship that at standard temperature and pressure (0°C and 1 atmosphere), one mole of gas occupies a volume of 22.4 liters.

3. e. *Students know* how to calculate the masses of reactants and products in a chemical reaction from the mass of one of the reactants or products and the relevant atomic masses.

Atoms are neither created nor destroyed in a chemical reaction. When the chemical reaction is written as a balanced expression, it is possible to calculate the mass of any one of the products or of any one of the reactants if the mass of just one reactant or product is known.

Students can be taught how to use balanced chemical equations to predict the mass of any product or reactant. Teachers should emphasize that the coefficients in the balanced chemical equation are mole quantities, not masses. Here is an example: How many grams of water will be obtained by combining 5.0 grams of

hydrogen gas with an excess of oxygen gas, according to the following balanced equation?

$$2H_2 + O_2 \rightarrow 2H_2O$$

This calculation is often set up algebraically, for example, as

$$5\text{ g } H_2 \times \frac{1\text{ mole } H_2}{2\text{g } H_2} \times \frac{2\text{ mole } H_2O}{2\text{ mole } H_2} \times \frac{18\text{ g } H_2O}{1\text{ mole } H_2O} = ?\text{ g } H_2O$$

and can be easily completed by direct calculation and unit cancellation (dimensional analysis). Students should learn to recognize that the coefficients in the balanced equations refer to moles rather than to mass.

3. f.* *Students know* how to calculate percent yield in a chemical reaction.

Students can use a balanced equation for a chemical reaction to calculate from the given masses of reactants the masses of the resulting products. The masses so calculated represent the theoretical 100 percent conversion of reactants to products. When one or more of the products are weighed, the masses are often less than the theoretical 100 percent yield. One explanation is that the reaction may not go to completion and therefore will not convert all the reactants to products. A second possible explanation is that product material is lost in the separation and purification process. A third reason is that alternative reactions are taking place, leading to products different from those predicted. Percent yield is a standard way to compare actual and theoretical yields. It is defined as

$$\frac{\text{actual yield of product obtained in an experiment}}{\text{theoretical yield as determined by stoichiometric calculations}} \times 100\%$$

3. g.* *Students know* how to identify reactions that involve oxidation and reduction and how to balance oxidation-reduction reactions.

Oxidation of an element is defined as an increase in oxidation number, or a loss of electrons. *Reduction of an element* is defined as a decrease in oxidation number, or a gain in electrons. The assignment of oxidation numbers, or oxidation states, is a bookkeeping device. This process may be defined as the charge assigned to an atom, as though all the electrons in each bond were located on the more electronegative atom in the bond. Students should be taught how to assign oxidation numbers to atoms in free elements and in compounds. Lists of rules for this procedure are commonly found in chemistry textbooks.

In many important chemical reactions, elements change their oxidation states. These changes are called *redox,* or *oxidation-reduction reactions.* Respiration and photosynthesis are common examples with which students are familiar.

Any chemical reaction in which electrons are transferred from one substance to another is an oxidation-reduction reaction. Transfer can be determined by checking the oxidation states of atoms in reactants and products. A single displacement

reaction, such as zinc metal in a copper sulfate solution, is a typical redox reaction because the zinc atoms lose two electrons and the copper ions gain two electrons. Redox reactions may be balanced by ensuring that the number of electrons lost in one part of the reaction equals the number of electrons gained in another part of the reaction. This principle and the conservation of atoms will hold true in any balanced oxidation/reduction reaction.

Students should be taught how to use the half-reaction method for balancing redox equations. This method divides the reaction into an oxidation portion and a reduction portion, both of which are balanced separately and then combined into an overall equation with no net change in number of electrons.

STANDARD SET 4. Gases and Their Properties

Standard Set 4 requires a knowledge of applicable physical concepts and sufficient skills in mathematical problem solving to describe (1) gases at the molecular level; (2) the behavior of gases; and (3) the measurable properties of gases. As background, students should know the physical states of matter and their properties and know how mixtures, especially homogeneous mixtures, differ from pure substances. Students should know that temperature measures how hot a system is, regardless of the system's size, and that heat flows from a region of higher temperature to one of lower temperature. Familiarity with the Celsius and Kelvin temperature scales is necessary. Knowledge of the motion of particles (and objects) and of kinetic energy is also required. Students should have the mathematical background to recognize and use directly and inversely proportional relationships. The ability to solve algebraic equations with several given quantities and one unknown is essential.

The main assumptions of the kinetic molecular theory should be covered in presenting this material, and a connection should be made to the behavior of ideal gases. Students are often more comfortable with studying the volume and temperature of a gas, but the concept of pressure also needs to be dealt with thoroughly. With knowledge of the kinetic molecular theory, students should see how motions and collisions of particles produce measurable properties, such as pressure.

4. The kinetic molecular theory describes the motion of atoms and molecules and explains the properties of gases. As a basis for understanding this concept:

a. *Students know* the random motion of molecules and their collisions with a surface create the observable pressure on that surface.

Fluids consist of molecules that freely move past each other in random directions. Intermolecular forces hold the atoms or molecules in liquids close to each other. Gases consist of tiny particles, either atoms or molecules, spaced far apart from each other and reasonably free to move at high speeds, near the speed of sound. In the study of chemistry, gases and liquids are considered fluids.

Pressure is defined as force per unit area. The force in fluids comes from collisions of atoms or molecules with the walls of the container. Air pressure is created by the weight of the gas in the atmosphere striking surfaces. Gravity pulls air molecules toward Earth, the surface that they strike. Water pressure can be understood in the same fashion, but the pressures are much greater because of the greater density of water. Pressure in water increases with depth, and pressure in air decreases with altitude. However, pressure is felt equally in all directions in fluids because of the random motion of the molecules.

4. b. *Students know* the random motion of molecules explains the diffusion of gases.

Another result of the kinetic molecular theory is that gases diffuse into each other to form homogeneous mixtures. An excellent demonstration of diffusion is the white ammonium chloride ring formed by simultaneous diffusion of ammonia vapor and hydrogen chloride gas toward the middle of a glass tube. The white ring forms nearer the region where hydrogen chloride was introduced, illustrating both diffusion and the principle that heavier gases have a slower rate of diffusion.

4. c. *Students know* how to apply the gas laws to relations between the pressure, temperature, and volume of any amount of an ideal gas or any mixture of ideal gases.

A fixed number of moles *n* of gas can have different values for pressure *P,* volume *V,* and temperature *T.* Relationships among these properties are defined for an ideal gas and can be used to predict the effects of changing one or more of these properties and solving for unknown quantities. Students should know and be able to use the three gas law relationships summarized in Table 1, "Gas Law Relationships."

Table 1
Gas Law Relationships

Expression of gas laws	*Fixed values*	*Variable relationships*	*Form for calculations*
PV = constant	*n, T*	Inverse	$P_1V_1 = P_2V_2$
V/T = constant	*n, P*	Direct	$V_1/T_1 = V_2/T_2$
P/T = constant	*n, V*	Direct	$P_1/T_1 = P_2/T_2$

The first expression of the gas law shown in Table 1 is sometimes taught as Boyle's law and the second as Charles's law, according to the historical order of their discovery. They are both simpler cases of the more general ideal gas law introduced in Standard 4.h in this section. For a fixed number of moles of gas, a combined gas law has the form PV/T = constant, or $P_1V_1/T_1 = P_2V_2/T_2$. This law is useful in calculations where *P, V,* and *T* are changing. By placing a balloon over the

mouth of an Erlenmeyer flask, the teacher can demonstrate that volume divided by temperature equals a constant. When the flask is heated, the balloon inflates; when the flask is cooled, the balloon deflates.

4. d. *Students know* the values and meanings of standard temperature and pressure (STP).

Standard temperature is 0°C, and standard pressure (STP) is 1 atmosphere (760 mm Hg). These standards are an agreed-on set of conditions for gases against which to consider other temperatures and pressures. When volumes of gases are being compared, the temperature and pressure must be specified. For a fixed mass of gas at a specified temperature and pressure, the volume is also fixed.

4. e. *Students know* how to convert between the Celsius and Kelvin temperature scales.

Some chemical calculations require an absolute temperature scale, called the Kelvin scale (K), for which the coldest possible temperature is equal to zero. There are no negative temperatures on the Kelvin scale. In theory if a sample of any material is cooled as much as possible, the lowest temperature that can be reached is 0 K, experimentally determined as equivalent to −273.15°C. The Kelvin scale starts with absolute zero (0 K) because of this theoretical lowest temperature limit. A Kelvin temperature is always 273.15 degrees greater than an equivalent Celsius temperature, but a Kelvin temperature is specified without the degree symbol. The magnitude of one unit of change in the K scale is equal to the magnitude of one unit of change on the °C scale.

4. f. *Students know* there is no temperature lower than 0 Kelvin.

The kinetic molecular theory is the basis for understanding heat and temperature. The greater the atomic and molecular motion, the greater the observed temperature of a substance. If all atomic and molecular motion stopped, the temperature of the material would reach an absolute minimum. This minimum is absolute zero, or −273.15°C. The third law of thermodynamics states that this temperature can never be reached. Experimental efforts to create very low temperatures have resulted in lowering the temperature of objects to within a fraction of a degree of absolute zero.

4. g.* *Students know* the kinetic theory of gases relates the absolute temperature of a gas to the average kinetic energy of its molecules or atoms.

The value of the average kinetic energy for an ideal gas is directly proportional to its Kelvin temperature. Average kinetic energy can be related to changes in pressure and volume as a function of temperature. At 0 K all motion in an ideal monatomic gas ceases, meaning that the average kinetic energy equals zero.

4. h.* *Students know* how to solve problems by using the ideal gas law in the form $PV = nRT$.

The relationships among pressure, volume, and temperature for a fixed mass of gas can be expressed as the ideal gas law, $PV = nRT$, where n represents moles and R represents the universal gas constant, which is 0.0821 liter-atmosphere per mole-Kelvin (this unit can be abbreviated as (L atm)/(mol K) or as L atm mol^{-1} K^{-1}).

4. i. * *Students know* how to apply Dalton's law of partial pressures to describe the composition of gases and Graham's law to predict diffusion of gases.

It is important to distinguish clearly between *diffusion* and *effusion*. *Diffusion* is the process by which separate atoms or molecules intermingle as a result of random motion. *Effusion* is the process by which gas molecules pass from one container to another at lower pressure through a very small opening.

Graham's law states that the rates of effusion of two gases at the same temperature and pressure are inversely proportional to the square roots of their molar masses. Graham's law also approximately applies to rates of diffusion for gases, although diffusion is a more complicated process to describe than effusion.

Dalton's law of partial pressures states that total pressure in a gas-filled container is equal to the sum of the partial pressures of the component gases. This law can be introduced by showing that ideal gases have properties based solely on the number of moles present in the sample, without regard to the chemical identities of the gas particles involved.

STANDARD SET 5. Acids and Bases

Students who learn the concepts in this standard set will be able to understand and explain aqueous acid–base reactions, properties of acids and bases, and pH as a measure of acidity and basicity. Careful thought should be given to how this standard set best fits with the other chemistry standards. It may be desirable to cover Standard Set 6, "Solutions," in this section first so that students will know about the aqueous dissolving process and about concentration calculations and units. Familiarity with Standard Set 9, "Chemical Equilibrium," in this section may help students to conceptualize the strengths of acids and bases.

Students should be able to represent and balance chemical reactions. They should also be able to interpret periodic trends in electronegativity for the upper two rows of the periodic table (see Standard Set 1, "Atomic and Molecular Structure," in this section). Hydrogen with its low electronegativity easily forms the positive hydrogen ion, H^+. Students need to know the charge and formula of the hydroxide ion, OH^-. With knowledge of polar covalent bonding, students should

be able to distinguish between two important types of neutral molecular compounds: those that dissolve in an aqueous solution and remain almost completely as neutral molecules and those that dissolve in an aqueous solution and partially or almost completely produce charged ions (see Standard Sets 1 and 2 in this section).

Students should be able to compare the three descriptions of acids and bases—the Arrhenius, Brønsted-Lowry, and Lewis acid–base definitions—and recognize electron lone pairs on Lewis dot structures of molecules (see Standard Set 2, "Chemical Bonds," in this section). To calculate pH, students should understand and be able to use base-10 logarithms and antilogarithms and know how to obtain logarithms by using a calculator. Students should become proficient at converting between pH, pOH, $[H^+]$, and $[OH^-]$.

5. Acids, bases, and salts are three classes of compounds that form ions in water solutions. As a basis for understanding this concept:

a. *Students know* the observable properties of acids, bases, and salt solutions.

Comparing and contrasting the properties of acids and bases provide a context for understanding their behavior. Some observable properties of acids are that they taste sour; change the color of litmus paper from blue to red; indicate acidic values on universal indicator paper; react with certain metals to produce hydrogen gas; and react with metal hydroxides, or bases, to produce water and a salt. Some observable properties of bases are that basic substances taste bitter or feel slippery; change the color of litmus paper from red to blue; indicate basic values on universal indicator paper; and react with many compounds containing hydrogen ions, or acids, to produce water and a salt.

These properties can be effectively demonstrated by using extracted pigment from red cabbage as an indicator to analyze solutions of household ammonia and white vinegar at various concentrations. When the indicator is added, basic solutions turn green, and acidic solutions turn red. Students can also use universal indicator solutions to test common household substances. **Students need to follow established safety procedures while conducting experiments.**

5. b. *Students know* acids are hydrogen-ion-donating and bases are hydrogen-ion-accepting substances.

According to the Brønsted-Lowry acid–base definition, acids donate hydrogen ions, and bases accept hydrogen ions. Acids that are formed from the nonmetals found in the first and second rows of the periodic table easily dissociate to produce hydrogen ions because these nonmetals have a large electronegativity compared with that of hydrogen. Once students know that acids and bases have different effects on the same indicator, they are ready to deepen their understanding of acid–base behavior at the molecular level. Examples and studies of chemical reactions should be used to demonstrate these definitions of acids and bases.

5. c. *Students know* strong acids and bases fully dissociate and weak acids and bases partially dissociate.

Acids dissociate by donating hydrogen ions, and bases ionize by dissociating to form hydroxide ions (from a hydroxide salt) or by accepting hydrogen ions. Some acids and bases either dissociate or ionize almost completely, and others do so only partially. Nearly complete dissociation is strong; partial dissociation is weak. The strength of an acid or a base can vary, depending on such conditions as temperature and concentration.

5. d. *Students know* how to use the pH scale to characterize acid and base solutions.

The pH scale measures the concentrations of hydrogen ions in solution and the acidic or basic nature of the solution. The scale is not linear but logarithmic, meaning that at pH 2, for example, the concentration of hydrogen ions is ten times greater than it is at pH 3. The pH scale ranges from below 0 (very acidic) to above 14 (very basic). Students should learn that pH values less than 7 are considered acidic and those greater than 7 are considered basic.

5. e.* *Students know* the Arrhenius, Brønsted-Lowry, and Lewis acid–base definitions.

Other acid–base definitions beside Brønsted-Lowry (see Standard 5.b in this section) are the Arrhenius and Lewis definitions. An Arrhenius base must contain hydroxide, such as in the chemical KOH. NH_3 does not contain hydroxide and therefore would not be a base according to the Arrhenius definition. However, because NH_3 accepts hydrogen ions, it would be a base according to the Brønsted-Lowry definition. A Lewis acid is an electron pair receptor, and a Lewis base is an electron pair donor. Using the Lewis definition extends the concept of acid–base reactions to nonaqueous systems. The compound BF_3, for example, is an acid because it accepts a lone electron pair, but BF_3 would not be an acid according to the Brønsted-Lowry definition because the hydrogen ion is not present.

5. f.* *Students know* how to calculate pH from the hydrogen-ion concentration.

The pH scale is defined as $-\log_{10}[H^+]$, where $[H^+]$ is the hydrogen-ion concentration in moles per liter of solution. Students should be taught to convert between hydrogen-ion concentration and pH.

5. g.* *Students know* buffers stabilize pH in acid–base reactions.

A buffer is a solution that stabilizes H^+ concentration levels. Such a solution may release hydrogen ions as pH rises or consume hydrogen ions as pH decreases. A critically important, but extremely complex example is the equilibria between

carbon dioxide, carbonic acid, bicarbonate, carbonate, and solid calcium carbonate that keeps the world's oceans at a nearly constant pH of about 8. Another measure students should learn is pOH, which is the negative logarithm of the OH^- concentration expressed in moles per liter of solution. The sum of pH and pOH is always 14.0 for a given solution at 25°C.

STANDARD SET 6. Solutions

As background for this standard set, students must know the physical states of matter and the corresponding properties. Specifically, the properties of the liquid state are important because most discussions will focus on liquid solutions. Students also need an understanding of molecules and ions and sufficient mathematical skills (especially an understanding of unit conversions) to describe solutions at the ionic or molecular level, describe the dissolution process, and quantify solution concentrations. Moreover, students should be able to calculate mass and volume in a variety of units and work with ratios, percentages, and moles (see Standard Set 3, "Conservation of Matter and Stoichiometry").

The sequencing of instruction regarding these standards deserves special attention and consideration. The knowledge of mixtures (especially of homogeneous mixtures), which will be acquired from the study of these standards, can be helpful in reinforcing the concepts of electronegativity, covalent bonding, shapes of molecules, and ionic compounds (see Standard Set 1, "Atomic and Molecular Structure," and Standard Set 2, "Chemical Bonds"). Some of these standards may also be helpful prerequisite knowledge for understanding Standard Set 5, "Acids and Bases." These standards also help prepare students to study the separation or purification of solution components as covered in Standard Set 9, "Chemical Equilibrium." Toward that end, instruction in these standards should include the concept of solubility as an equilibrium between solid and solute forms of a substance and the study of reactions that lead to precipitate formation.

6. Solutions are homogeneous mixtures of two or more substances. As a basis for understanding this concept:

a. *Students know* the definitions of *solute* and *solvent*.

Simple solutions are homogeneous mixtures of two substances. A *solute* is the dissolved substance in a solution, and a *solvent* is, by quantity, the major component in the solution.

6. b. *Students know* how to describe the dissolving process at the molecular level by using the concept of random molecular motion.

The kinetic molecular theory as applied to gases can be extended to explain how the solute and solvent particles are in constant random motion. The kinetic energy

of this motion causes diffusion of the solute into the solvent, resulting in a homogeneous solution. When a solid is in contact with a liquid, at least some small degree of dissolution always occurs. The equilibrium concentration of solute in solvent will depend on the surface interactions between the molecules of solute and solvent. Equilibrium is reached when all competing processes are in balance. Those processes include the tendency for dissolved molecules to spread randomly in the solvent and the competing strength of the bonds and other forces among solute molecules, among solvent molecules, and between solute and solvent molecules. When salts dissolve in water, positive and negative ions are separated and surrounded by polar water molecules.

6. c. *Students know* temperature, pressure, and surface area affect the dissolving process.

In a liquid solvent, solubility of gases and solids is a function of temperature. Students should have experience with reactions in which precipitates are formed or gases are released from solution, and they should be taught that the concentration of a substance that appears as solid or gas must exceed the solubility of the solvent.

Increasing the temperature usually increases the solubility of solid solutes but always decreases the solubility of gaseous solutes. An example of a solid ionic solute compound that decreases in solubility as the temperature increases is Na_2SO_4. An example of one that increases in solubility as the temperature increases is $NaNO_3$. The solubility of a gas in a liquid is directly proportional to the pressure of that gas above the solution. It is important to distinguish solubility equilibrium from rates of dissolution. Concepts of equilibrium describe only how much solute will dissolve at equilibrium, not how quickly this process will occur.

6. d. *Students know* how to calculate the concentration of a solute in terms of grams per liter, molarity, parts per million, and percent composition.

All concentration units listed previously are a measure of the amount of solute compared with the amount of solution. Grams per liter represent the mass of solute divided by the volume of solution. *Molarity* describes moles of solute divided by liters of solution. Students can calculate the number of moles of dissolved solute and divide by the volume in liters of the total solution, yielding units of moles per liter. Parts per million, which is a ratio of one part of solute to one million parts of solvent, is usually applied to very dilute solutions. Percent composition is the ratio of one part of solvent to one hundred parts of solvent and is expressed as a percent. To calculate parts per million and percent composition, students determine the mass of solvent and solute and then divide the mass of the solute by the total mass of the solution. This number is then multiplied by 10^6 and expressed as parts per million (ppm) or by 100 and expressed as a percent.

6. e.* *Students know* the relationship between the molality of a solute in a solution and the solution's depressed freezing point or elevated boiling point.

The physical properties of the freezing point and boiling point of a solution are directly proportional to the concentration of the solution in molality. *Molality* is similar to *molarity* except that molality expresses moles of solute dissolved in a kilogram of solvent rather than in a liter of solution. In other words, *molality* is the amount of solute present divided by the amount of solvent present. Molality is a convenient measure because it does not depend on volume and therefore does not change with temperature. Sometimes physical properties change with concentration of solute; for example, salt, such as sodium chloride or calcium chloride, is sprinkled on icy roads to lower the freezing point of water and melt the snow or ice. The freezing point is lowered, or depressed, in proportion to the amount of salt dissolved.

6. f.* *Students know* how molecules in a solution are separated or purified by the methods of chromatography and distillation.

Chromatography is a powerful, commonly used method to separate substances for analysis, including DNA, protein, and metal ions. The principle takes advantage of a moving solvent and a stationary substrate to induce separation. A useful, interesting example is paper chromatography. In this laboratory technique a mixture of solutes, such as ink dyes, is applied to a sheet of chromatographic paper. One end of the paper is dipped in a solvent that moves (wicks) up or along the paper. Solutes (the various ink dyes, for example) separate into bands of colors. Solutes with great affinity for the paper move little, those with less move more, and those with very little affinity may travel with the leading edge of the solvent.

Mixtures can sometimes be separated by distillation, which capitalizes on differences in the forces holding molecules in a liquid state. Crude oil, for example, is processed by commercial refineries (by a catalytic reaction called "cracking") and separated by heat distillation to give a variety of commercial products, from highly volatile kerosene and gasoline to heavier oils used to lubricate engines or to heat homes.

STANDARD SET 7. Chemical Thermodynamics

Students should know the relationship between heat and temperature and understand the concepts of kinetic energy and motion at the molecular level. They should also know how to differentiate between chemical and physical change and understand that both types of change entail the loss of free energy, a process that results in increased stability. The standards marked with an asterisk in this

section require students to know the concepts presented in the physics section of this chapter, Standard Set 3, "Heat and Thermodynamics." Those standards provide a foundation for understanding the kinetic molecular model and introduce students to concepts of heat and entropy.

7. Energy is exchanged or transformed in all chemical reactions and physical changes of matter. As a basis for understanding this concept:

a. *Students know* how to describe temperature and heat flow in terms of the motion of molecules (or atoms).

Temperature is a measure of the average kinetic energy of molecular motion in a sample. *Heat* is energy transferred from a sample at higher temperature to one at lower temperature. Often, heat is described as flowing from the system to the surroundings or from the surroundings to the system. The system is defined by its boundaries, and the surroundings are outside the boundaries, with "the universe" frequently considered as the surroundings.

7. b. *Students know* chemical processes can either release (exothermic) or absorb (endothermic) thermal energy.

Endothermic processes absorb heat, and their equations can be written with heat as a reactant. *Exothermic processes* release heat, and their equations can be written with heat as a product. The net heat released to or absorbed from the surroundings comes from the making and breaking of chemical bonds during a reaction. Students understand and relate heat to the internal motion of the atoms and molecules. They also understand that breaking a bond always requires energy and that making a bond almost always releases energy. The amount of energy per bond depends on the strength of the bond.

The potential energy of the reaction system may be plotted for the different reaction stages: reactants, transition states, and products. This plot will show reactants at lower potential energy than products for an endothermic reaction and reactants at higher potential energy than products for an exothermic reaction. A higher energy transition state usually exists between the reactant and product energy states that affect the reaction rate covered in Standard Set 8, "Reaction Rates," in this section.

7. c. *Students know* energy is released when a material condenses or freezes and is absorbed when a material evaporates or melts.

Physical changes are accompanied by changes in internal energy. Changes of physical state either absorb or release heat. Evaporation and melting require energy to overcome the bonds of attractions in the corresponding liquid or solid state. Condensation and freezing release heat to the surroundings as internal energy is reduced and bonds of attraction are formed.

7. d. *Students know* how to solve problems involving heat flow and temperature changes, using known values of specific heat and latent heat of phase change.

Qualitative knowledge that students gained by mastering the previous standards will help them to solve problems related to the heating or cooling of a substance over a given temperature range. *Specific heat* is the energy needed to change the temperature of one gram of substance by one degree Celsius. The unit of specific heat is joule/gram-degree.

During phase changes, energy is added or removed without a corresponding temperature change. This phenomenon is called *latent* (or *hidden*) *heat.* There is a latent heat of fusion and a latent heat of vaporization. The unit of latent heat is joule/gram or kilojoule/mole. Students should be able to diagram the temperature changes that occur when ice at a temperature below zero is heated to superheated steam, which has temperatures above 100°C.

7. e.* *Students know* how to apply Hess's law to calculate enthalpy change in a reaction.

As samples of elements combine to make compounds, heat may be absorbed from the environment. If one mole of a compound is formed from elements and all substances begin and end at 25°C and are under standard atmospheric pressure, the heat absorbed during the compound's synthesis is known as its *standard enthalpy of formation,* H_f. If heat is not absorbed but released, H_f is negative. Values for the enthalpy of formation of thousands of compounds are available in reference books.

Hess's law states that if a chemical reaction is carried out in any imaginable series of steps, the net enthalpy change (heat absorbed) in the reaction is the sum of the enthalpy changes for the individual steps. For example, a reaction can be imagined to proceed in just two steps: first, making its reactants into elements, and second, making those elements into products. The enthalpy change in a reaction $a\mathrm{A} + b\mathrm{B} = c\mathrm{C} + d\mathrm{D}$ is the quantity ΔH_r^o defined as "the heat absorbed when a moles of chemical A and b moles of B react to make c moles of C and d moles of D." Through the use of Hess's law and the two steps of the example, the enthalpy change for this reaction can be expressed in terms of the enthalpy of formation H_f^o of each of the chemicals: the enthalpy change is the (weighted) sum of the heat of formation for each of the products minus the sum for each of the reactants, or

$$\Delta H_r^o = [c\,H_f^o(\mathrm{C}) + d\,H_f^o(\mathrm{D})] - [a\,H_f^o(\mathrm{A}) + b\,H_f^o(\mathrm{B})]$$

By using the balanced equation for a chemical reaction and the enthalpy of formation for each chemical, students can calculate the heat absorbed or released when a given quantity of a reactant is consumed.

7. f.* *Students know* how to use the Gibbs free energy equation to determine whether a reaction would be spontaneous.

Endothermic and exothermic reactions can be spontaneous under standard conditions of temperature and pressure. Therefore, releasing heat and going to a lower energy state cannot be the only force driving chemical reactions. The tendency to disorder, or entropy, is the other driving force. A convenient conceptual way to account for the balance between these two driving forces, enthalpy changes and changes in disorder, was developed by J. Willard Gibbs and is called the Gibbs free-energy change ΔG.

In the Gibbs free-energy equation (shown below), ΔH is the change in enthalpy, T is the Kelvin temperature, and ΔS is the change in entropy. (Note that students may need to be introduced to the basics of entropy, as presented in Standard Set 3 for physics in this chapter).

$$\Delta G = \Delta H - T\Delta S$$

The Gibbs free-energy change is used to predict in which direction a reaction will proceed. A negative value for the Gibbs free-energy change predicts the formation of products (a spontaneous reaction); a positive value predicts, or favors, reactants (a nonspontaneous reaction). Standard values for the Gibbs free energy of elements and compounds at a specified temperature are available in tables.

STANDARD SET 8. Reaction Rates

To describe rates of chemical reactions, factors affecting rates, and the energy changes involved, students need to know that chemical reactions consume reactants and form products (see Standard Set 3, "Conservation of Matter and Stoichiometry," in this section). Students should be able to explain chemical reactions at the molecular level and know how kinetic energy at the molecular level is measured by temperature (see Standard Sets 3 and 7 in this section). Students have acquired knowledge of pressure and volume relationships for gases (see Standard Set 4, "Gases and Their Properties," in this section) and can plot potential energy versus course of reaction for endothermic and exothermic reactions (see Standard Set 7, "Chemical Thermodynamics," in this section). The ability to calculate rates of change from slopes of lines and curves is required.

8. Chemical reaction rates depend on factors that influence the frequency of collision of reactant molecules. As a basis for understanding this concept:

a. *Students know* the rate of reaction is the decrease in concentration of reactants or the increase in concentration of products with time.

Students may have an intuitive idea that reaction rate is a measure of how fast reactions proceed, but a quantitative measure for reaction rate also is needed. For

example, explosive reactions are very fast, as are many biological reactions in the cell; and other reactions, such as iron rusting, are very slow. *Reaction rate* is defined as the rate of decrease in concentration of reactants or as the rate of increase in concentration of products, and these reciprocal changes form a balanced equation that reflects the conservation of matter. Students can see from the balanced equation that as the reaction proceeds, the concentration of reactants must decrease, and the concentration of products must increase in proportion to their mole ratios.

8. b. *Students know* how reaction rates depend on such factors as concentration, temperature, and pressure.

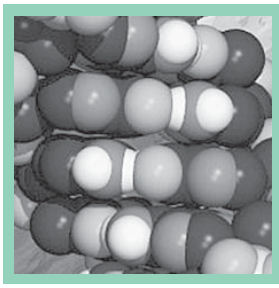
Concentration, temperature, and pressure should be emphasized because they are major factors affecting the collision of reactant molecules and, thus, affecting reaction rates. Increasing the concentration of reactants increases the number of collisions per unit time. Increasing temperature (which increases the average kinetic energy of molecules) also increases the number of collisions per unit time. Though the collision rate modestly increases, the greater kinetic energy dramatically increases the chances of each collision leading to a reaction (e.g., the Arrhenius effect). Increasing pressure increases the reaction rate only when one or more of the reactants or products are gases. With gaseous reactants, increasing pressure is the same as increasing concentration and results in an elevated reaction rate.

8. c. *Students know* the role a catalyst plays in increasing the reaction rate.

A *catalyst* increases the rate of a chemical reaction without taking part in the net reaction. A catalyst lowers the energy barrier between reactants and products by promoting a more favorable pathway for the reaction. Surfaces often play important roles as catalysts for many reactions. One reactant might be temporarily held on the surface of a catalyst. There the bonds of the reactant may be weakened, allowing another substance to react with it more quickly. Living systems speed up life-dependent reactions with biological catalysts called *enzymes.* Catalysts are used in automobile exhaust systems to reduce the emission of smog-producing unburned hydrocarbons.

8. d.* *Students know* the definition and role of activation energy in a chemical reaction.

Even in a spontaneous reaction, reactants are usually required to pass through a transition state that has a higher energy than either the reactants or the products. The additional energy, called the *activation energy,* or the *activation barrier,* is related to such factors as strength of bonding within the reactants. The more energy required to go from reactants to activated transition complex, the higher the activation barrier, and the slower a reaction will be. Catalysts speed up rates by lowering the activation barrier along the reaction pathway between products and reactants.



STANDARD SET 9. Chemical Equilibrium

Chemical equilibrium is a dynamic state. To understand the factors affecting equilibrium and to write expressions used to quantify a state of equilibrium, students will need a thorough knowledge of reaction rates (see Standard Set 8, "Reaction Rates," in this section) and of chemical thermodynamics (see Standard Set 7, "Chemical Thermodynamics," in this section). Changes in heat accompanying chemical reactions and spontaneity of chemical reactions are key topics. Students should be able to identify the physical states of substances undergoing chemical reactions and use knowledge of substances at the atomic and molecular levels (see Standard Set 3, "Conservation of Matter and Stoichiometry," in this section).

Students need to know how gases respond to changes in pressure and volume (see Standard Set 4, "Gases and Their Properties," in this section). Students need to be able to calculate concentration and molarity for solutions, particularly for aqueous solutions (see Standard Set 6, "Solutions," in this section). Students familiar with acid–base reactions (see Standard Set 5, "Acids and Bases," in this section) and precipitation reactions (see Standard Set 6 in this section) will have an advantage in learning the concept of equilibrium.

To calculate equilibrium constants, students should be able to balance chemical equations (see Standard Set 3 in this section), work readily with concentration and pressure units (see Standard Sets 4 and 6 in this section), and use exponents in mathematical calculations.

Students often have difficulty in understanding that equilibrium, which is a dynamic process, occurs when no net changes in a product or reactant concentration take place. An analogy can be made with a pair of escalators operating between two floors. If the same number of people go up as go down in a ten-minute interval, the rate of people moving up equals the rate of people moving down. Overall, any extra people arriving on one floor are canceled out by others leaving the floor. Therefore, the number of people on each floor will be constant over time, and the population of the two floors is in dynamic equilibrium. This analogy can be extended to a chemical reaction by considering that if the number of moles produced in one direction of the reaction is the same as the number consumed in the opposite direction, then the reaction has reached a state of dynamic equilibrium. Students will learn that when a stress is applied to a chemical reaction in equilibrium, a shift will occur to partly relieve the stress.

9. Chemical equilibrium is a dynamic process at the molecular level. As a basis for understanding this concept:

a. *Students know* how to use Le Chatelier's principle to predict the effect of changes in concentration, temperature, and pressure.

Le Chatelier's principle can be introduced by emphasizing the balanced nature of an equilibrium system. If an equilibrium system is stressed or disturbed, the

system will respond (change or shift) to partially relieve or undo the stress. A new equilibrium will eventually be established with a new set of conditions. When the stress is applied, the reaction is no longer at equilibrium and will shift to regain equilibrium. For instance, if the concentration of a reactant in a system in dynamic equilibrium is decreased, products will be consumed to produce more of that reactant. Students need to remember that heat is a reactant in endothermic reactions and a product in exothermic reactions. Therefore, increasing temperature will shift an endothermic reaction, for example, to the right to regain equilibrium. Students should note that any endothermic chemical reaction is exothermic in the reverse direction.

Pressure is proportional to concentration for gases; therefore, for chemical reactions that have a gaseous product or reactant, pressure affects the system as a whole. Increased pressure shifts the equilibrium toward the smaller number of moles of gas, alleviating the pressure stress. If both sides of the equilibrium have an equal number of moles of gas, increasing pressure does not affect the equilibrium. Adding an inert gas, such as argon, to a reaction will not change the partial pressures of the reactant or product gases and therefore will have no effect on the equilibrium.

9. b. *Students know* equilibrium is established when forward and reverse reaction rates are equal.

Forward and reverse reactions at equilibrium are going on at the same time and at the same rate, causing overall concentrations of each reactant and product to remain constant over time.

9. c.* *Students know* how to write and calculate an equilibrium constant expression for a reaction.

Because the concentrations of substances in a system at chemical equilibrium are constant over time, chemical expressions related to each concentration will also be constant. Here is a general equation for a reaction at equilibrium:

$$a\mathrm{A} + b\mathrm{B} \leftrightarrow c\mathrm{C} + d\mathrm{D}$$

The general expression for the equilibrium constant of a chemical reaction is K_{eq}, defined at a particular temperature, often 25°C. Its formula is

$$K_{eq} = \frac{[\mathrm{C}]^c[\mathrm{D}]^d}{[\mathrm{A}]^a[\mathrm{B}]^b}$$

When K_{eq} is being calculated, only gaseous substances and aqueous solutions are considered. Equilibrium concentrations of products, in moles per liter, are in the numerator, and equilibrium concentrations of reactants are in the denominator. The exponents are the corresponding coefficients from the balanced chemical equation. A large K_{eq} means the forward reaction goes almost to completion; that is, little reverse reaction occurs. A very small K_{eq} means the reverse reaction goes almost to completion, or little forward reaction occurs. The solubility product constant K_{sp} is the equilibrium constant for salts in solution.

STANDARD SET 10. Organic and Biochemistry

A solid understanding of chemical and biological concepts is required to describe the versatility with which carbon atoms form molecules and to illustrate the structure of organic and biological molecules and the polymers they form. An understanding of these concepts is also necessary to be able to name organic molecules and to identify organic functional groups. Students also need to know the material in atomic and molecular structure (see Standard Set 1, "Atomic and Molecular Structure," in this section), especially as the concepts pertain to carbon and nearby nonmetals and to electronegativity.

Chemical bonds (see Standard Set 2, "Chemical Bonds," in this section), especially the topics of covalent bonding and Lewis dot structures, should also be well understood. Students should know the importance of biochemical compounds—proteins, carbohydrates, and nucleic acids—and understand their roles in living organisms. They must be familiar with single, double, and triple bonds to name organic molecules correctly (see Standard Set 2 in this section) and with the covalent bonding and electronegativity of nitrogen and oxygen to identify functional groups of organic molecules and to understand amino acids and proteins (see Standard Sets 1 and 2 in this section).

10. The bonding characteristics of carbon allow the formation of many different organic molecules of varied sizes, shapes, and chemical properties and provide the biochemical basis of life. As a basis for understanding this concept:

a. *Students know* large molecules (polymers), such as proteins, nucleic acids, and starch, are formed by repetitive combinations of simple subunits.

Students can readily visualize large molecules called polymers as consisting of repetitive and systematic combinations of smaller, simpler groups of atoms, including carbon. All polymeric molecules, including biological molecules, such as proteins, nucleic acids, and starch, are made up of various unique combinations of a relatively small number of chemically simple subunits. For example, starch is a polymer made from a large number of simple sugar molecules joined together.

10. b. *Students know* the bonding characteristics of carbon that result in the formation of a large variety of structures ranging from simple hydrocarbons to complex polymers and biological molecules.

Building on what they learned in grade eight about the unique bonding characteristics of carbon, students explore in greater depth the incredible diversity of carbon-based molecules. They are reminded that, given carbon's four bonding electrons and four vacancies available to form bonds, carbon is able to form stable

covalent bonds—single or multiple—with other carbon atoms and with atoms of other elements.

Students learn how the presence of single, double, and triple bonds determines the geometry of carbon-based molecules. The variety of these molecules is enormous: over 16 million carbon-containing compounds are known. The compounds range from simple hydrocarbon molecules (e.g., methane and ethane) to complex organic polymers and biological molecules (e.g., proteins) and include many manufactured polymers used in daily life (e.g., polyester, nylon, and polyethylene).

10. c. *Students know* amino acids are the building blocks of proteins.

Proteins are large single-stranded polymers often made up of thousands of relatively small subunits called *amino acids.* The bond attaching two amino acids, known as the *peptide bond,* is identical for any pair of amino acids. The chemical composition of the amino acid itself varies. Variation in composition and ordering of amino acids gives protein molecules their unique properties and shapes. These properties and shapes define the protein's functions, many of which are essential to the life of an organism. The blueprint for building the protein molecules is deoxyribonucleic acid (DNA). Biotechnology is advancing rapidly as more is learned about DNA, amino acid sequences, and the shapes and functions of proteins.

10. d.* *Students know* the system for naming the ten simplest linear hydrocarbons and isomers that contain single bonds, simple hydrocarbons with double and triple bonds, and simple molecules that contain a benzene ring.

Organic molecules can be simple or extremely complex. The naming system for these molecules, however, is relatively straightforward and reflects the composition and structure of each molecule. Each name is made up of a prefix and a suffix. The prefix tells the number of carbon atoms in the longest continuous sequence of the molecule, and the suffix indicates the kind of bond between carbon atoms. For example, the four simplest hydrocarbon molecules are methane, ethane, propane, and butane. The prefixes, *meth-, eth-, prop-,* and *but-* refer to one, two, three, and four carbons, respectively. The *-ane* ending indicates that there are only carbon-carbon single bonds. *Ene* endings are used for double bonds and *-yne* for triple bonds. Benzene, C_6H_6, is a flat hexagonally shaped molecule of six carbon atoms bonded to each other. Many compounds can be built by substitutions on straight-chain hydrocarbons and benzene rings.

10. e.* *Students know* how to identify the functional groups that form the basis of alcohols, ketones, ethers, amines, esters, aldehydes, and organic acids.

Organic molecules are grouped into classes based on patterns of bonding between carbon and noncarbon atoms (e.g., nitrogen and oxygen). Groups based on unique patterns of bonding are called *functional groups.* Examples of these groups

are alcohols, ketones, ethers, amines, esters, aldehydes, and organic (carboxylic) acids.

10. f.* *Students know* the R-group structure of amino acids and know how they combine to form the polypeptide backbone structure of proteins.

Amino acid molecules have a well-known structure, and all contain a side chain called an *R-group.* Differences in the R-group are the basis for differences between the amino acids. Bonding two amino acids creates a *dipeptide,* bonding three creates a *tripeptide,* and adding more creates a polymer called a *polypeptide.* Polypeptides made biologically are called *proteins.*

STANDARD SET 11. Nuclear Processes

This section requires a knowledge of chemical and physical concepts and sufficient mathematical skills to describe the nucleus and its subatomic particles. Topics covered are nuclear reactions and their accompanying changes in energy and forms of radiation and quantification of radioactive decay as a function of time. Students should know about nuclear structure and properties, the mass and charge of the proton and neutron, and the use of the periodic table to determine the number of protons in an atom's nucleus (see Standard Set 1, "Atomic and Molecular Structure," in this section). Students should also know how to calculate and use percentages to determine the amount of radioactive substance remaining after a time interval of disintegration.

Students should be introduced to this standard set with a review of the nucleus and its constituent protons and neutrons. Simple hydrogen, deuterium, and tritium nuclei can be used to introduce and define *isotopes.* Students should be reminded of the difference between an element's average atomic mass and the mass number for a specific isotope. They should already know that electrons and protons attract each other as do any particles of opposite charge, but they probably do not know what holds the protons and neutrons in the nucleus together. Teachers should introduce students to the strong nuclear force and explain how it holds protons and neutrons together and how it can overcome the repulsion between charged protons at very close distances. Students should also be introduced to quarks as the constituents of protons and neutrons.

11. Nuclear processes are those in which an atomic nucleus changes, including radioactive decay of naturally occurring and human-made isotopes, nuclear fission, and nuclear fusion. As a basis for understanding this concept:

a. *Students know* protons and neutrons in the nucleus are held together by nuclear forces that overcome the electromagnetic repulsion between the protons.

The nucleus is held together by the strong nuclear force. The strong nuclear force acts between protons, between neutrons, and between protons and neutrons but has a limited range comparable to the size of an atomic nucleus. The nuclear force is able to overcome the mutual electrostatic repulsion of the protons only when the protons and neutrons are near each other as they are in the nucleus of an atom.

11. b. *Students know* the energy release per gram of material is much larger in nuclear fusion or fission reactions than in chemical reactions. The change in mass (calculated by $E = mc^2$) is small but significant in nuclear reactions.

Two major types of nuclear reactions are fusion and fission. In *fusion* reactions two nuclei come together and merge to form a heavier nucleus. In *fission* a heavy nucleus splits apart to form two (or more) lighter nuclei. The binding energy of a nucleus depends on the number of neutrons and protons it contains. A general term for a proton or a neutron is a *nucleon.* In both fusion and fission reactions, the total number of nucleons does not change, but large amounts of energy are released as nucleons combine into different arrangements. This energy is one million times more than energies involved in chemical reactions.

11. c. *Students know* some naturally occurring isotopes of elements are radioactive, as are isotopes formed in nuclear reactions.

Sometimes atoms with the same number of protons in the nucleus have different numbers of neutrons. These atoms are called *isotopes* of an element. Both naturally occurring and human-made isotopes of elements can be either stable or unstable. Less stable isotopes of one element, called *parent isotopes,* will undergo radioactive decay, transforming to more stable isotopes of another element, called *daughter products,* which can also be either stable or radioactive. For a radioactive isotope to be found in nature, it must either have a long half-life, such as potassium-40, uranium-238, uranium-235, or thorium-232, or be the daughter product, such as radon-222, of a parent with a long half-life, such as uranium-238.

11. d. *Students know* the three most common forms of radioactive decay (alpha, beta, and gamma) and know how the nucleus changes in each type of decay.

Radioactive isotopes transform to more stable isotopes, emitting particles from the nucleus. These particles are helium-4 nuclei (alpha radiation), electrons or positrons (beta radiation), or high-energy electromagnetic rays (gamma radiation). Isotopes of elements that undergo alpha decay produce other isotopes with two less protons and two less neutrons than the original isotope. Uranium-238, for instance, emits an alpha particle and becomes thorium-234.

Isotopes of elements that undergo beta decay produce elements with the same number of nucleons but with one more proton or one less proton. For example, thorium-234 beta decays to protactinium-234, which then beta decays to uranium-234. Alpha and beta decay are ionizing radiations with the potential to damage surrounding materials. After alpha and beta decay, the resulting nuclei often emit high-energy photons called *gamma rays.* This process does not change the number of nucleons in the nucleus of the isotope but brings about a lower energy state in the nucleus.

11. e. *Students know* alpha, beta, and gamma radiation produce different amounts and kinds of damage in matter and have different penetrations.

Alpha, beta, and gamma rays are *ionizing radiations,* meaning that those rays produce tracks of ions of atoms and molecules when they interact with materials. For all three types of rays, the energies of particles emitted in radioactive decay are typically for each particle on the order of 1MeV, equal to 1.6×10^{-13} joule, which is enough energy to ionize as many as half a million atoms.

Alpha particles have the shortest ranges, and matter that is only a few millimeters thick will stop them. They will not penetrate a thick sheet of paper but will deposit all their energy along a relatively short path, resulting in a high degree of ionization along that path.

Beta particles have longer ranges, typically penetrating matter up to several centimeters thick. Those particles are electrons or positrons (the antimatter electron), have one unit of either negative or positive electric charge, and are approximately 1/2000 of the mass of a proton. These high-energy electrons have longer ranges than alpha particles and deposit their energy along longer paths, spreading the ionization over a greater distance in the material.

Gamma rays can penetrate matter up to several meters thick. Gamma rays are high-energy photons that have no electric charge and no rest mass (the structural energy of the particle). They will travel unimpeded through materials until they strike an electron or the nucleus of an atom. The gamma ray's energy will then be either completely or partially absorbed, and neighboring atoms will be ionized. Therefore, these three types of radiation interact with matter by losing energy and ionizing surrounding atoms.

Alpha radiation is dangerous if ingested or inhaled. For example, radon-222, a noble gas element, is a naturally occurring hazard in some regions. Living organisms or sensitive materials can be protected from ionizing radiation by shielding them and increasing their distance from radiation sources.

Because many people deeply fear and misunderstand radioactivity, chemistry teachers should address and explore the ability of each form of radiation to penetrate matter and cause damage. Students may be familiar with radon detection devices, similar to smoke detectors, found in many homes. Discussion of biological and health effects of ionizing radiation can inform students about the risks and benefits of nuclear reactions. Videos can be used in the classroom to show demonstrations of the penetrating ability of alpha, beta, and gamma radiation through paper, aluminum, and lead or through other dense substances of varying thicknesses. Geiger counter measurements can be used to record radiation data. The order of penetrating ability, from greatest to least, is gamma > beta > alpha, and this order is the basis for assessing proper shielding of radiation sources for safety.

There are a number of naturally occurring sources of ionizing radiation. One is potassium-40, which can be detected easily in potash fertilizer by using a Geiger counter. The other is background cosmic and alpha radiation from radon. This radiation can be seen in cloud chambers improvised in the classroom.

11. f.* *Students know* how to calculate the amount of a radioactive substance remaining after an integral number of half-lives have passed.

Radioactive decay transforms the initial (parent) nuclei into more stable (daughter) nuclei with a characteristic half-life. The *half-life* is the time it takes for one-half of a given number of parent atoms to decay to daughter atoms. One-half of the remaining parent atoms will then decay to produce more daughter atoms in the next half-life period. It is possible to predict only the proportion, not the individual number, of parent atoms that will undergo decay. Therefore, after one half-life, 50 percent of the initial parent nuclei remain; after two half-lives, 25 percent; and so forth. The intensity of radiation from a radioactive source is related to the half-life and to the original number of radioactive atoms present.

11. g.* *Students know* protons and neutrons have substructures and consist of particles called *quarks.*

Just as atoms consist of subatomic nuclear particles (protons and neutrons), so do protons and neutrons have constituent particles called *quarks.* Quarks come in six different types, but only two types are involved in ordinary matter: the *up quark* and the *down quark.*

For enrichment beyond the content in the standard, the following description is included: Protons and neutrons each contain three quarks. A proton consists of two up quarks and one down quark. A neutron consists of two down quarks and one up quark. Quarks have fractional electric charges; therefore, the charge of an

up quark is $+\frac{2}{3}$ units, and the charge of a down quark is $-\frac{1}{3}$ units. It is believed that it is not possible to isolate a free quark. All the common matter in the material world is made up of just three fundamental particles: the up quark, the down quark, and the electron.

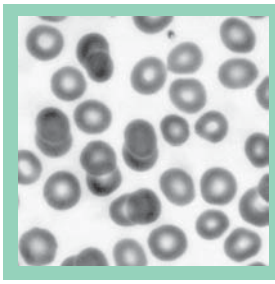
Biology/Life Sciences

Living organisms appear in many variations, yet there are basic similarities among their forms and functions. For example, all organisms require an outside source of energy to sustain life processes; all organisms demonstrate patterns of growth and, in many cases, senescence, the process of becoming old; and the continuity of all species requires reproduction. All organisms are constructed from the same types of macromolecules (proteins, nucleic acids, lipids) and inherit a deoxyribonucleic acid (DNA) genome from a parent or parents. DNA is always transcribed to yield ribonucleic acid (RNA), which is translated through the use of a nearly universal genetic code. Environmental factors frequently regulate and influence the expression of specific genes.

Biologists study life at many levels, and the biology standards for grades nine through twelve reflect these studies. *Organisms* are part of an ecosystem and have complex relationships with other organisms and the physical environment. Ecologists study these populations and communities, and many are deeply interested in the physical and behavioral adaptations of organisms. Evolutionary biologists share these interests because the fitness of an organism is a manifestation of these adaptations. *Adaptations* are traits subject to the rules of inheritance; therefore, genetics and evolutionary biology are closely allied fields.

Physiologists study whole body systems or organs. For example, a neurophysiologist focuses primarily on the nervous system. Cell biologists study the details of how cells and organelles work, considering such weighty matters as how cytoskeletal elements segregate chromosomes during mitosis, how proteins are sorted to different compartments of the cell, and how receptors in the cell membrane communicate with factors that regulate gene expression. Many cell biologists also consider themselves to be developmental biologists, molecular geneticists, or biochemists. There are many connections between all the fields and different ways of viewing life.

Biology textbooks typically start with a review of chemistry and energetics; therefore, California students will be able to make good use of their study of the content standards for "Chemistry of Living Systems" in the eighth grade. The principles of cellular biology, including respiration and photosynthesis, are usually taught next, followed by instruction in molecular and Mendelian genetics. Population genetics and evolution follow naturally from the study of genetics and lead to a discussion of diversity of form and physiology. The teaching culminates with ecology, a subject that draws on each of the preceding topics. The teaching comes full circle because ecology is also a starting point for students in lower elementary school grade levels.



STANDARD SET 1. Cell Biology

The first knowledge of cells came from the work of an English scientist, Robert Hooke, who in 1665 used a primitive microscope to study thin sections of cork and called the boxlike cavities he saw "cells." Antony van Leeuwenhoek later observed one-celled "animalcules" in pond water, but not until the 1830s did Theodor Schwann view cartilage tissue in which he discovered cells resembling plant cells. He published the theory that cells are the basic unit of life. Rudolf Virchow used the work of Schwann and Matthias Schleiden to advance the cell theory, presenting the concept that plants and animals are made of cells that contain fluid and nuclei and arise from preexisting cells.

After the cell theory was established, detailed study of cell structure and function depended on the improvement of microscopes and on techniques for preparing specimens for observation. It is now understood that cells in plants and animals contain genes to control chemical reactions needed for survival and organelles to perform those reactions. Living organisms may consist of one cell, as in bacteria, or of many cells acting in a coordinated and cooperative manner, as in plants, animals, and fungi. All cells have at least three structures in common: genetic material, a cell or plasma membrane, and cytoplasm.

1. The fundamental life processes of plants and animals depend on a variety of chemical reactions that occur in specialized areas of the organism's cells. As a basis for understanding this concept:

a. *Students know* cells are enclosed within semipermeable membranes that regulate their interaction with their surroundings.

The plasma membrane consists of two layers of lipid molecules organized with the polar (globular) heads of the molecules forming the outside of the membrane and the nonpolar (straight) tails forming the interior of the membrane. Protein molecules embedded within the membrane move about relative to one another in a fluid fashion. Because of its dynamic nature the membrane is sometimes referred to as *the fluid mosaic model* of membrane structure.

Cell membranes have three major ways of taking in or of regulating the passage of materials into and out of the cell: simple diffusion, carrier-facilitated diffusion, and active transport. Osmosis of water is a form of diffusion. Simple diffusion and carrier-facilitated diffusion do not require the expenditure of chemical bond energy, and the net movement of materials reflects a concentration gradient or a voltage gradient or both. Active transport requires free energy, in the form of either chemical bond energy or a coupled concentration gradient, and permits the net transport or "pumping" of materials against a concentration gradient.

1. b. *Students know* enzymes are proteins that catalyze biochemical reactions without altering the reaction equilibrium and the activities of enzymes depend on the temperature, ionic conditions, and the pH of the surroundings.

Almost all *enzymes* are protein catalysts made by living organisms. Enzymes speed up favorable (spontaneous) reactions by reducing the activation energy required for the reaction, but they are not consumed in the reactions they promote. To demonstrate the action of enzymes on a substrate, the teacher can use liver homogenate or yeast as a source of the enzyme catalase and hydrogen peroxide as the substrate. The effect of various environmental factors, such as pH, temperature, and substrate concentration, on the rate of reaction can be investigated. These investigations should encourage student observation, recording of qualitative and quantitative data, and graphing and interpretation of data.

1. c. *Students know* how prokaryotic cells, eukaryotic cells (including those from plants and animals), and viruses differ in complexity and general structure.

All living cells are divided into one of two groups according to their cellular structure. *Prokaryotes* have no membrane-bound organelles and are represented by the Kingdom Monera, which in modern nomenclature is subdivided into the Eubacteria and Archaea. *Eukaryotes* have a complex internal structure that allows thousands of chemical reactions to proceed simultaneously in various organelles. *Viruses* are not cells; they consist of only a protein coat surrounding a strand of genetic material, either RNA or DNA.

1. d. *Students know* the central dogma of molecular biology outlines the flow of information from transcription of ribonucleic acid (RNA) in the nucleus to translation of proteins on ribosomes in the cytoplasm.

DNA, which is found in the nucleus of eukaryotes, contains the genetic information for encoding proteins. The DNA sequence specifying a specific protein is copied (transcribed) into messenger RNA (mRNA), which then carries this message out of the nucleus to the ribosomes located in the cytoplasm. The mRNA message is then translated, or converted, into the protein originally coded for by the DNA.

1. e. *Students know* the role of the endoplasmic reticulum and Golgi apparatus in the secretion of proteins.

There are two types—rough and smooth—of endoplasmic reticulum (ER), both of which are systems of folded sacs and interconnected channels. *Rough ER* synthesizes proteins, and *smooth ER* modifies or detoxifies lipids. Rough ER produces new proteins, including membrane proteins. The proteins to be exported from the cell are moved to the Golgi apparatus for modification, packaged in vesicles, and transported to the plasma membrane for secretion.

1. f. *Students know* usable energy is captured from sunlight by chloroplasts and is stored through the synthesis of sugar from carbon dioxide.

Photosynthesis is a complex process in which visible sunlight is converted into chemical energy in carbohydrate molecules. This process occurs within chloroplasts and specifically within the thylakoid membrane (light-dependent reaction) and the stroma (light-independent reaction). During the light-dependent reaction, water is oxidized and light energy is converted into chemical bond energy generating ATP, NADPH + H^+, and oxygen gas.[†] During the light-independent reaction (Calvin cycle), carbon dioxide, ATP, and NADPH + H^+ react, forming phosphoglyceraldehyde, which is then converted into sugars. By using a microscope with appropriate magnification, students can see the chloroplasts in plant cells (e.g., lettuce, onion) and photosynthetic protists (e.g., euglena).

Students can prepare slides of these cells themselves, an activity that provides a good opportunity to see the necessity for well-made thin sections of specimens and for correct staining procedures. Commercially prepared slides are also available. By observing prepared cross sections of a leaf under a microscope, students can see how a leaf is organized structurally and think about the access of cells to light and carbon dioxide during photosynthesis. The production of oxygen from photosynthesis can be demonstrated and measured quantitatively with a volumeter, which can collect oxygen gas from the illuminated leaves of an aquatic plant, such as elodea. By varying the distance between the light source and the plant, teachers can demonstrate intensities of the effects of various illumination. To eliminate heat as a factor, the teacher can place a heat sink, such as a flat-sided bottle of water, between the plant and light source to absorb or dissipate unwanted heat.

1. g. *Students know* the role of the mitochondria in making stored chemical-bond energy available to cells by completing the breakdown of glucose to carbon dioxide.

Mitochondria consist of a matrix where three-carbon fragments originating from carbohydrates are broken down (to CO_2 and water) and of the cristae where ATP is produced. Cell respiration occurs in a series of reactions in which fats, proteins, and carbohydrates, mostly glucose, are broken down to produce carbon dioxide, water, and energy. Most of the energy from cell respiration is converted into ATP, a substance that powers most cell activities.

1. h. *Students know* most macromolecules (polysaccharides, nucleic acids, proteins, lipids) in cells and organisms are synthesized from a small collection of simple precursors.

Many of the large carbon compound molecules necessary for life (e.g., polysaccharides, nucleic acids, proteins, and lipids) are polymers of smaller monomers. Polysaccharides are composed of monosaccharides; proteins are composed of amino

[†]ATP is adenosine triphosphate, and NADPH is reduced nicotinamide adenine dinucleotide phosphate.

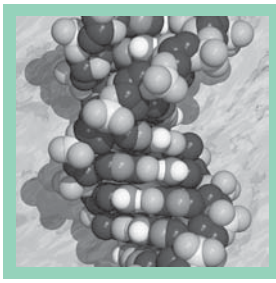
acids; lipids are composed of fatty acids, glycerol, and other components; and nucleic acids are composed of nucleotides.

1. i.* *Students know* how chemiosmotic gradients in the mitochondria and chloroplast store energy for ATP production.

Enzymes called *ATP synthase,* located within the thylakoid membranes in chloroplasts and cristae membranes in mitochondria, synthesize most ATP within cells. The thylakoid and cristae membranes are impermeable to protons except at pores that are coupled with the ATP synthase. The potential energy of the proton concentration gradient drives ATP synthesis as the protons move through the ATP synthase pores. The proton gradient is established by energy furnished by a flow of electrons passing through the electron transport system located within these membranes.

1. j.* *Students know* how eukaryotic cells are given shape and internal organization by a cytoskeleton or cell wall or both.

The *cytoskeleton,* which gives shape to and organizes eukaryotic cells, is composed of fine protein threads called *microfilaments* and thin protein tubes called *microtubules. Cilia* and *flagella* are composed of microtubules arranged in the 9 + 2 arrangement, in which nine pairs of microtubules surround two single microtubules. The rapid assembly and disassembly of microtubules and microfilaments and their capacity to slide past one another enable cells to move, as observed in white blood cells and amoebae, and also account for movement of organelles within the cell. Students can observe prepared slides of plant mitosis in an onion root tip to see the microtubules that make up the spindle apparatus. Prepared slides of white fish blastula reveal animal spindle apparatus and centrioles, both of which are composed of microtubules.



STANDARD SET 2. Genetics (Meiosis and Fertilization)

Students should know that organisms reproduce offspring of their own kind and that organisms of the same species resemble each other. Students have been introduced to the idea that some characteristics can be passed from parents to offspring and that individual variations appear among offspring and in the broader population. Understanding genetic variation requires mastery of the fundamentals of sex cell formation and the steps to reorganize and redistribute genetic material during defined stages in the cell cycle.

Students should understand the difference between asexual cell reproduction (*mitosis*) and the formation of male or female gamete cells (*meiosis*). Sexual reproduction initially requires the production of haploid eggs and haploid sperm, a process occurring in humans within the female ovary and the male testis. These haploid cells unite in fertilization and produce the *diploid zygote,* or fertilized cell.

The mechanisms involved in synapsis and movement of chromosomes during meiosis bring about the halving of the chromosome numbers for the production of the haploid male or female gamete cells from the original diploid parent cell and different combinations of parental genes. The exchange of chromosomal segments between homologous chromosomes (crossing over) revises the association of genes on the chromosomes and contributes to increased diversity. Any change in genetic constitution through mutation, crossing over, or chromosome assortment during meiosis promotes genetic variation in a population.

2. Mutation and sexual reproduction lead to genetic variation in a population. As a basis for understanding this concept:

a. *Students know* meiosis is an early step in sexual reproduction in which the pairs of chromosomes separate and segregate randomly during cell division to produce gametes containing one chromosome of each type.

Haploid gamete production through meiosis involves two cell divisions. During meiosis prophase I, the homologous chromosomes are paired, a process that abets the exchange of chromosome parts through breakage and reunion. The second meiotic division parallels the mechanics of mitosis except that this division is not preceded by a round of DNA replication; therefore, the cells end up with the haploid number of chromosomes. (The nucleus in a haploid cell contains one set of chromosomes.) Four haploid nuclei are produced from the two divisions that characterize meiosis, and each of the four resulting cells has different chromosomal constituents (components). In the male all four become sperm cells. In the female only one becomes an egg, while the other three remain small degenerate polar bodies and cannot be fertilized.

Chromosome models can be constructed and used to illustrate the segregation taking place during the phases of mitosis (covered initially in Standard 1.e for grade seven in Chapter 4) and meiosis. Commercially available optical microscope slides also show cells captured in mitosis (onion root tip) or meiosis (Ascaris blastocyst cells), and computer and video animations are also available.

2. b. *Students know* only certain cells in a multicellular organism undergo meiosis.

Only special diploid cells, called *spermatogonia* in the testis of the male and *oogonia* in the female ovary, undergo meiotic divisions to produce the haploid sperm and haploid eggs.

2. c. *Students know* how random chromosome segregation explains the probability that a particular allele will be in a gamete.

The steps in meiosis involve random chromosome segregation, a process that accounts for the probability that a particular allele will be packaged in any given gamete. This process allows for genetic predictions based on laws of probability

that pertain to genetic sortings. Students can create a genetic chart and mark alternate traits on chromosomes, one expression coming from the mother and another expression coming from the father. Students can be shown that partitions of the chromosomes are controlled by chance (are random) and that separation (segregation) of chromosomes during *karyokinesis* (division of the nucleus) leads to the random sequestering of different combinations of chromosomes.

2. d. *Students know* new combinations of alleles may be generated in a zygote through the fusion of male and female gametes (fertilization).

Once gametes are formed, the second half of sexual reproduction can take place. In this process a diploid organism is reconstituted from two haploid parts. When a sperm is coupled with an egg, a fertilized egg (zygote) is produced that contains the combined genotypes of the parents to produce a new allelic composition for the progeny. Genetic charts can be used to illustrate how new combinations of alleles may be present in a zygote through the events of meiosis and the chance union of gametes. Students should be able to read the genetic diploid karyotype, or chromosomal makeup, of a fertilized egg and compare the allelic composition of progeny with the genotypes and phenotypes of the parents.

2. e. *Students know* why approximately half of an individual's DNA sequence comes from each parent.

Chromosomes are composed of a single, very long molecule of double-stranded DNA and proteins. Genes are defined as segments of DNA that code for polypeptides (proteins). During fertilization half the DNA of the progeny comes from the gamete of one parent, and the other half comes from the gamete of the other parent.

2. f. *Students know* the role of chromosomes in determining an individual's sex.

The normal human somatic cell contains 46 chromosomes, of which 44 are pairs of homologous chromosomes and 2 are sex chromosomes. Females usually carry two X chromosomes, and males possess one X and a smaller Y chromosome. Combinations of these two sex chromosomes determine the sex of the progeny.

2. g. *Students know* how to predict possible combinations of alleles in a zygote from the genetic makeup of the parents.

When the genetic makeups of potential parents are known, the possible assortments of alleles in their gametes can be determined for each genetic locus. Two parental gametes will fuse during fertilization, and with all pair-wise combinations of their gametes considered, the possible genetic makeups of progeny can then be predicted.

STANDARD SET 3. Genetics (Mendel's Laws)

Breeding of plants and animals has been an active technology for thousands of years, but the science of heredity is linked to the genetics pioneer Gregor Mendel. He studied phenotypic traits of various plants, especially those of peas. (A *phenotypic trait* is the physical appearance of a trait in an organism). From the appearance of these traits in different generations of growth, he was able to infer their *genotypes* (the genetic makeup of an organism with respect to a trait) and to speculate about the genetic makeup and method of transfer of the hereditary units from one generation to the next. (Probability analysis is now used to predict probable progeny phenotypes from various parental genetic crosses.) The genetic basis for Mendel's laws of segregation and independent assortment is apparent from genetic outcomes of crosses.

3. A multicellular organism develops from a single zygote, and its phenotype depends on its genotype, which is established at fertilization. As a basis for understanding this concept:

a. *Students know* how to predict the probable outcome of phenotypes in a genetic cross from the genotypes of the parents and mode of inheritance (autosomal or X-linked, dominant or recessive).

Monohybrid crosses, including autosomal dominant alleles, autosomal recessive alleles, incomplete dominant alleles, and X-linked alleles, can be used to indicate the parental genotypes and phenotypes. The possible gametes derived from each parent are based on genotypic ratios and can be used to predict possible progeny. The predictive (probabilistic) methods for determining the outcome of genotypes and phenotypes in a genetic cross can be introduced by using Punnett Squares and probability mathematics.

Teachers should review the process of writing genotypes and help students translate genotypes into phenotypes. Teachers should emphasize dominant, recessive, and incomplete dominance as the students advance to an explanation of monohybrid crosses illustrating human conditions characterized by autosomal recessive alleles, such as albinism, cystic fibrosis, Tay-Sachs, and phenylketonuria (PKU). These disorders can be contrasted with those produced by possession of just one autosomal dominant allele, conditions such as Huntington disease, dwarfism, and neurofibromatosis. This basic introduction can be followed with examples of incomplete dominance, such as seen in the comparisons of straight, curly, and wavy hair or in the expression of intermediate flower colors in snapdragon plants.

Sex-linked characteristics that are found only on the X chromosome should also be considered, and students should reflect on how this mode of transmission can cause the exclusive or near-exclusive appearance in males of color blindness, hemophilia, fragile-X syndrome, and sex-linked muscular dystrophy.

3. b. *Students know* the genetic basis for Mendel's laws of segregation and independent assortment.

Mendel deduced that for each characteristic, an organism inherits two genes, one from each parent. When the two alleles differ, the dominant allele is expressed, and the recessive allele remains hidden. Two genes or alleles separate (segregate) during gamete production in meiosis, resulting in the sorting of alleles into separate gametes (the law of segregation). Students can be shown how to diagram Mendel's explanation for how a trait present in the parental generation can appear to vanish in the first filial (F1) generation of a monohybrid cross and then reappear in the following second filial (F2) generation.

Students should be told that alternate versions of a gene at a single locus are called *alleles.* Students should understand Mendel's deduction that for each character, an organism inherits two genes, one from each parent. From this point students should realize that if the two alleles differ, the dominant allele, if there is one, is expressed, and the recessive allele remains hidden. Students should recall that the two genes, or alleles, separate (segregate) during gamete production in meiosis and that this sorting of alleles into separate gametes is the basis for the law of segregation. This law applies most accurately when genes reside on separate chromosomes that segregate out at random, and it often does not apply or is a poor predictor for combinations and frequencies of genes that reside on the same chromosome. Students can study various resources that describe Mendel's logic and build models to illustrate the laws of segregation and independent assortment.

3. c.* *Students know* how to predict the probable mode of inheritance from a pedigree diagram showing phenotypes.

Students should be taught how to use a pedigree diagram showing phenotypes to predict the mode of inheritance.

3. d.* *Students know* how to use data on frequency of recombination at meiosis to estimate genetic distances between loci and to interpret genetic maps of chromosomes.

Students should be able to interpret genetic maps of chromosomes and manipulate genetic data by using standard techniques to relate recombination at meiosis to estimate genetic distances between loci.

STANDARD SET 4. Genetics (Molecular Biology)

All cells contain DNA as their genetic material. The role of DNA in organisms is twofold: first, to store and transfer genetic information from one generation to the next, and second, to express that genetic information in the synthesis

of proteins. By controlling protein synthesis, DNA controls the structure and function of all cells. The complexity of an organism determines whether it may have several hundred to more than twenty thousand proteins as a part of its makeup.

Proteins are composed of a sequence of amino acids linked by peptide bonds (see Standard 10.c for chemistry in this chapter). The identity, number, and sequence of the amino acids in a protein give each protein its unique structure and function. Twenty types of amino acids are commonly employed in proteins, and each can appear many times in a single protein molecule. The proper sequence of amino acids in a protein is translated from an RNA sequence that is itself encoded in the DNA.

4. Genes are a set of instructions encoded in the DNA sequence of each organism that specify the sequence of amino acids in proteins characteristic of that organism. As a basis for understanding this concept:

a. *Students know* the general pathway by which ribosomes synthesize proteins, using tRNAs to translate genetic information in mRNA.

DNA does not leave the cell nucleus, but messenger RNA (mRNA), complementary to DNA, carries encoded information from DNA to the ribosomes (transcription) in the cytoplasm. (The *ribosomes* translate mRNAs to make protein.) Freely floating amino acids within the cytoplasm are bonded to specific transfer RNAs (tRNAs) that then transport the amino acid to the mRNA now located on the ribosome. As a ribosome moves along the mRNA strand, each mRNA codon, or sequence of three nucleotides specifying the insertion of a particular amino acid, is paired in sequence with the anticodon of the tRNA that recognizes the sequence. Each amino acid is added, in turn, to the growing polypeptide at the specified position.

After learning about transcription and translation through careful study of expository texts, students can simulate the processes on paper or with representative models. Computer software and commercial videos are available that illustrate animated sequences of transcription and translation.

4. b. *Students know* how to apply the genetic coding rules to predict the sequence of amino acids from a sequence of codons in RNA.

The sequence of amino acids in protein is provided by the genetic information found in DNA. In prokaryotes, mRNA transcripts of a coding sequence are copied from the DNA as a single contiguous sequence. In eukaryotes, the initial RNA transcript, while in the nucleus, is composed of *exons,* sequences of nucleotides that carry useful information for protein synthesis, and *introns,* sequences that do not. Before leaving the nucleus, the initial transcript is processed to remove introns and splice exons together. The processed transcript, then properly called mRNA and carrying the appropriate codon sequence for a protein, is transported from the nucleus to the ribosome for translation.

Each mRNA has sequences, called *codons,* that are decoded three nucleotides at a time. Each codon specifies the addition of a single amino acid to a growing polypeptide chain. A *start codon* signals the beginning of the sequence of codons to be translated, and a *stop codon* ends the sequence to be translated into protein. Students can write out mRNA sequences with start and stop codons from a given DNA sequence and use a table of the genetic code to predict the primary sequences of proteins.

4. c. *Students know* how mutations in the DNA sequence of a gene may or may not affect the expression of the gene or the sequence of amino acids in the encoded protein.

Mutations are permanent changes in the sequence of nitrogen-containing bases in DNA (see Standard 5.a in this section for details on DNA structure and nitrogen bases). Mutations occur when base pairs are incorrectly matched (e.g., *A* bonded to *C* rather than *A* bonded to *T*) and can, but usually do not, improve the product coded by the gene. Inserting or deleting base pairs in an existing gene can cause a mutation by changing the codon reading frame used by a ribosome. Mutations that occur in somatic, or nongerm, cells are often not detected because they cannot be passed on to offspring. They may, however, give rise to cancer or other undesirable cellular changes. Mutations in the germline can produce functionally different proteins that cause such genetic diseases as Tay-Sachs, sickle cell anemia, and Duchenne muscular dystrophy.

4. d. *Students know* specialization of cells in multicellular organisms is usually due to different patterns of gene expression rather than to differences of the genes themselves.

Gene expression is a process in which a gene codes for a product, usually a protein, through transcription and translation. Nearly all cells in an organism contain the same DNA, but each cell transcribes only that portion of DNA containing the genetic information for proteins required at that specific time by that specific cell. The remainder of the DNA is not expressed. Specific types of cells may produce specific proteins unique to that type of cell.

4. e. *Students know* proteins can differ from one another in the number and sequence of amino acids.

Protein molecules vary from about 50 to 3,000 amino acids in length. The types, sequences, and numbers of amino acids used determine the type of protein produced.

4. f.* *Students know* why proteins having different amino acid sequences typically have different shapes and chemical properties.

The 20 different protein-making amino acids have the same basic structure: an amino group; an acidic (carboxyl) group; and an R, or radical group (see Standard

Set 10, "Organic and Biochemistry," in the chemistry section of this chapter). The protein is formed by the amino group of one amino acid linking to the carboxyl group of another amino acid. This bond, called the *peptide bond,* is repeated to form long molecular chains with the R groups attached along the polymer backbone.

The properties of these amino acids vary from one another because of both the order and the chemical properties of these R groups. Typically, the long protein molecule folds on itself, creating a three-dimensional structure related to its function. Structure, for example, may allow a protein to be a highly specific catalyst, or enzyme, able to position and hold other molecules. The R group of an amino acid consists of atoms that may include carbon, hydrogen, nitrogen, oxygen, and sulfur, depending on the amino acid. Amino acids containing sulfur sometimes play an important role of cross-linking and stabilizing polymer chains. Because of their various R groups, different amino acids vary in their chemical and physical properties, such as solubility in water, electrical charge, and size. These differences are reflected in the unique structure and function of each type of protein.

STANDARD SET 5. Genetics (Biotechnology)

Long before scientists identified DNA as the genetic material of cells, much was known about inheritance and the relationships between various characteristics likely to appear from one generation to the next. However, to comprehend clearly how the genetic composition of cells changes, students need to understand the structure and activity of nucleic acids.

Genetic recombination occurs naturally in sexual reproduction, viral infection, and bacterial transformation. These natural events change the genetic makeup of organisms and provide the raw materials for evolution. Natural selection determines the usefulness of the recombinants. In recombinant DNA technology specific pieces of DNA are recombined in the laboratory to achieve a specific goal. The scientist, rather than natural selection, then determines the usefulness of the recombinant DNA created.

5. The genetic composition of cells can be altered by incorporation of exogenous DNA into the cells. As a basis for understanding this concept:

a. *Students know* the general structures and functions of DNA, RNA, and protein.

Nucleic acids are polymers composed of monomers called *nucleotides.* Each nucleotide consists of three subunits: a five-carbon pentose sugar, a phosphoric acid group, and one of four nitrogen bases. (For DNA these nitrogen bases are adenine, guanine, cytosine, or thymine.) DNA and RNA differ in a number of major ways. A DNA nucleotide contains a deoxyribose sugar, but RNA contains ribose sugar.

The nitrogen bases in RNA are the same as those in DNA except that thymine is replaced by uracil. RNA consists of only one strand of nucleotides instead of two as in DNA.

The DNA molecule consists of two strands twisted around each other into a double helix resembling a ladder twisted around its long axis. The outside, or uprights, of the ladder are formed by the two sugar-phosphate backbones. The rungs of the ladder are composed of pairs of nitrogen bases, one extending from each upright. In DNA these nitrogen bases always pair so that *T* pairs with *A,* and *G* pairs with *C.* This pairing is the reason DNA acts as a template for its own replication. RNA exists in many structural forms, many of which play different roles in protein synthesis. The mRNA form serves as a template during protein synthesis, and its codons are recognized by aminoacylated tRNAs. Protein and rRNA make up the structure of the ribosome.

Proteins are polymers composed of amino acid monomers (see Standard Set 10 for chemistry in this chapter). Different types of proteins function as enzymes and transport molecules, hormones, structural components of cells, and antibodies that fight infection. Most cells in an individual organism carry the same set of DNA instructions but do not use the entire DNA set all the time. Only a small amount of the DNA appropriate to the function of that cell is expressed. Genes are, therefore, turned on or turned off as needed by the cell, and the products coded by these genes are produced only when required.

5. b. *Students know* how to apply base-pairing rules to explain precise copying of DNA during semiconservative replication and transcription of information from DNA into mRNA.

Enzymes initiate DNA replication by unzipping, or unwinding, the double helix to separate the two parental strands. Each strand acts as a template to form a complementary daughter strand of DNA. The new daughter strands are formed when complementary new nucleotides are added to the bases of the nucleotides on the parental strands. The nucleotide sequence of the parental strand dictates the order of the nucleotides in the daughter strands. One parental strand is conserved and joins a newly synthesized complementary strand to form the new double helix; this process is called *semiconservative replication.*

DNA replication is usually initiated by the separation of DNA strands in a small region to make a "replication bubble" in which DNA synthesis is primed. The DNA strands progressively unwind and are replicated as the replication bubble expands, and the two forks of replication move in opposite directions along the chromosome. At each of the diverging replication forks, the strand that is conserved remains a single, continuous "leading" strand, and the other "lagging" complementary strand is made as a series of short fragments that are subsequently repaired and ligated together.

Students may visualize DNA by constructing models, and they can simulate semiconservative replication by tracing the synthesis of the leading and lagging

strands. The critical principles to teach with this activity are that two double-stranded DNA strands are the product of synthesis, that the process is semiconservative, that the antiparallel orientation of the strands requires repeated reinitiation on the lagging strand, and that the only information used during synthesis is specified by the base-pairing rules.

RNA is produced from DNA when a section of DNA (containing the nucleotide sequence required for the production of a specific protein) is transcribed. Only the template side of the DNA is copied. RNA then leaves the nucleus and travels to the cytoplasm, where protein synthesis takes place.

5. c. *Students know* how genetic engineering (biotechnology) is used to produce novel biomedical and agricultural products.

Recombinant DNA contains DNA from two or more different sources. Bacterial plasmids and viruses are the two most common vectors, or carriers, by which recombinant DNA is introduced into a host cell. Restriction enzymes provide the means by which researchers cut DNA at desired locations to provide DNA fragments with "sticky ends." Genes, once identified, can be amplified either by cloning or by polymerase chain reactions, both of which produce large numbers of copies. The recombinant cells are then grown in large fermentation vessels, and their products are extracted from the cells (or from the medium if the products are secreted) and purified. Genes for human insulin, human growth hormone, blood clotting factors, and many other products have been identified and introduced into bacteria or other microorganisms that are then cultured for commercial production. Some agricultural applications of this technology are the identification and insertion of genes to increase the productivity of food crops and animals and to promote resistance to certain pests and herbicides, robustness in the face of harsh environmental conditions, and resistance to various viruses.

Students can model the recombinant DNA process by using paper models to represent eukaryotic complementary DNA (cDNA), the activity of different restriction enzymes, and ligation into plasmid DNA containing an antibiotic resistance gene and origin of DNA replication. To manipulate the modeled DNA sequences, students can use scissors (representing the activity of restriction enzymes) and tape (representing DNA ligase). If both strands are modeled on a paper tape, students can visualize how, in many cases, restriction enzymes make staggered cuts that generate "sticky ends" and how the ends must be matched during ligation.

5. d.* *Students know* how basic DNA technology (restriction digestion by endonucleases, gel electrophoresis, ligation, and transformation) is used to construct recombinant DNA molecules.

In recombinant DNA technology DNA is isolated and exchanged between organisms to fulfill a specific human purpose. The desired gene is usually identified and extracted by using restriction enzymes, or *endonucleases,* to cut the DNA into fragments. Restriction enzymes typically cut *palindromic* portions of DNA, which

read the same forward and backward, in ways that form sticky complementary ends. DNA from different sources, but with complementary sticky ends, can be joined by the enzyme DNA ligase, thus forming recombinant DNA.

DNA fragments of varying lengths can be separated from one another by *gel electrophoresis.* In this process the particles, propelled by an electric current, are moved through an agarose gel. Depending on the size, shape, and electrical charge of the particles, they will move at different rates through the gel and thus form bands of particles of similar size and charge. With appropriate staining, the various DNA fragments can then be visualized and removed for further analysis or recombination.

5. e.* *Students know* how exogenous DNA can be inserted into bacterial cells to alter their genetic makeup and support expression of new protein products.

Bacteria can be induced to take up recombinant plasmids, a process called DNA transformation, and the plasmid is replicated as the bacteria reproduce. Recombinant bacteria can be grown to obtain billions of copies of the recombinant DNA. Commercially available kits containing all the necessary reagents, restriction enzymes, and bacteria are available for experiments in plasmid DNA transformation. Although the reagents and equipment can be expensive, various California corporations and universities have programs to make the cost more affordable, sometimes providing reagents and lending equipment.

Students should know that DNA transformation is a natural process and that horizontal DNA transfer is common in the wild. An example of how humans have manipulated genetic makeup is through the selective breeding of pets and of agricultural crops.

STANDARD SET 6. Ecology

Ecology is the study of relationships among living organisms and their interactions with the physical environment. These relationships are in a constant state of flux, and even small changes can cause effects throughout the ecosystem. Students in grades nine through twelve can be taught to think of ecology as changing relationships among the components of an ecosystem. Students also need to recognize that humans are participants in these ecosystem relationships, not just observers. A goal of classroom teaching should be to develop a strong scientific understanding of ecology to establish the basis for making informed and valid decisions.

6. Stability in an ecosystem is a balance between competing effects. As a basis for understanding this concept:

a. *Students know* biodiversity is the sum total of different kinds of organisms and is affected by alterations of habitats.

Biodiversity refers to the collective variety of living organisms in an ecosystem. This structure is influenced by alterations in habitat, including but not limited to climatic changes, fire, flood, and invasion by organisms from another system. The more biodiversity in an ecosystem, the greater its stability and resiliency. The best way for students to learn about ecology is to master the principles of the subject through careful study and then to make firsthand observations of ecosystems in action over time.

Although field trips are the ideal way to implement this process and should be encouraged, even career scientists often use models to study ecology. Local ecologists from government, private industry, or university programs may also be willing to serve as guest speakers in the classroom. Viewing the Internet's many virtual windows that show actual ecological experiments can also help students understand the scientific basis of ecology.

6. b. *Students know* how to analyze changes in an ecosystem resulting from changes in climate, human activity, introduction of nonnative species, or changes in population size.

Analysis of change can help people to describe and understand what is happening in a natural system and, to some extent, to control or influence that system. Understanding different kinds of change can help to improve predictions of what will happen next. Changes in ecosystems often manifest themselves in predictable patterns of climate, seasonal reproductive cycles, population cycles, and migrations. However, unexpected disturbances caused by human intervention or the introduction of a new species, for example, may destabilize the often complex and delicate balance in an ecosystem.

Analyzing changes in an ecosystem can require complex methods and techniques because variation is not necessarily simple and may be interrelated with changes or trends in other factors. Rates and patterns of change, including trends, cycles, and irregularities, are essential features of the living world and are useful indicators of change that can provide data for analysis. Often it is important to analyze change over time, a process called *longitudinal analysis.*

6. c. *Students know* how fluctuations in population size in an ecosystem are determined by the relative rates of birth, immigration, emigration, and death.

Fluctuations in the size of a population are often difficult to measure directly but may be estimated by measuring the relative rates of birth, death, immigration, and emigration in a population. The number of deaths and emigrations over time

will decrease a population's size, and the number of births and immigrations over time will increase it. Comparing rates for death and emigration with those for birth and immigration will determine whether the population shows a net growth or a decline over time.

6. d. *Students know* how water, carbon, and nitrogen cycle between abiotic resources and organic matter in the ecosystem and how oxygen cycles through photosynthesis and respiration.

Living things depend on nonliving things for life. At the organism level living things depend on natural resources, and at the molecular level, they depend on chemical cycles. Water, carbon, nitrogen, phosphorus, and other elements are recycled back and forth between organisms and their environments. Water, carbon, and nitrogen are necessary for life to exist. These chemicals are incorporated into plants (producers) by photosynthesis and nitrogen fixation and used by animals (consumers) for food and protein synthesis. Chemical recycling occurs through respiration, the excretion of waste products and, of course, the death of organisms.

6. e. *Students know* a vital part of an ecosystem is the stability of its producers and decomposers.

An ecosystem's producers (plants and photosynthetic microorganisms) and decomposers (fungi and microorganisms) are primarily responsible for the productivity and recycling of organic matter, respectively. Conditions that threaten the stability of producer and decomposer populations in an ecosystem jeopardize the availability of energy and the capability of matter to recycle in the rest of the biological community. To study the interaction between producers and decomposers, students can set up a closed or restricted ecosystem, such as a worm farm, a composting system, a terrarium, or an aquarium.

6. f. *Students know* at each link in a food web some energy is stored in newly made structures but much energy is dissipated into the environment as heat. This dissipation may be represented in an energy pyramid.

The energy pyramid illustrates how stored energy is passed from one organism to another. At every level in a food web, an organism uses energy metabolically to survive and grow, but much is released as heat, usually about 90 percent. At every link in a food web, energy is transferred to the next level, but typically only 10 percent of the energy from the previous level is passed on to the consumer.

6. g.* *Students know* how to distinguish between the accommodation of an individual organism to its environment and the gradual adaptation of a lineage of organisms through genetic change.

Living organisms may adapt to changing environments through nongenetic changes in their structure, metabolism, or behavior or through natural selection of

favorable combinations of alleles governing any or all of these processes. Genetic and behavioral adaptations are sometimes difficult to identify or to distinguish without studying the organism over a long time. Physical changes are slow to develop in most organisms, requiring careful measurements over many years. Examining fossil ancestors of an organism may help provide clues for detecting adaptation through genetic change. Genetic change can institute behavioral changes, making it all the more complicated to determine whether a change is solely a behavioral accommodation to environmental change.

Through the use of print and online resources in library-media centers, students can research the effects of encroaching urbanization on undeveloped land and consider the effects on specific species, such as the coyote (not endangered) and the California condor (endangered). Such examples can illustrate how some organisms adapt to their environments through learned changes in behavior, and others are unsuccessful in learning survival skills. Over a long time, organisms can also adapt to changing environments through genetic changes, some of which may include genetically determined changes in behavior. Such changes may be difficult to recognize because a long time must elapse before the changes become evident. Studies of the origins of desert pup fish or blind cave fish may help students understand how gradual genetic changes in an organism lead to adaptations to changes in its habitat.

STANDARD SET 7. Evolution (Population Genetics)

This discussion applies to Standard Sets 7 and 8. Students in grades nine through twelve should be ready to explore and understand the concept of biological evolution from its basis in genetics. The synthesis of genetics, and later of molecular biology, with the Darwin-Wallace theory of natural selection validated the mechanism of evolution and extended its scientific impact. Students need to understand that the same evolutionary mechanisms that have affected the rest of the living world have also affected the human species.

Students need to understand that a theory in science is not merely a hypothesis or a guess, but a unifying explanation of observed phenomena. Charles Darwin's theory of the origin of species by natural selection is such an explanation. Even though biologists continue to test the boundaries of this theory today, their investigations have not found credible evidence to refute the theory. Scientists have also had many opportunities to demonstrate the gradual evolution of populations in the wild and in controlled laboratory settings. As more populations of organisms are studied at the level of DNA sequence and as the fossil record improves, the understanding of species divergence has become clearer.

7. The frequency of an allele in a gene pool of a population depends on many factors and may be stable or unstable over time. As a basis for understanding this concept:

a. *Students know* why natural selection acts on the phenotype rather than the genotype of an organism.

Natural selection works directly on the expression or appearance of an inherited trait, the *phenotype,* rather than on the gene combination that produces that trait, the *genotype.* The influence of a dominant allele for a trait over a recessive one in the genotype determines the resulting phenotype on which natural selection acts.

7. b. *Students know* why alleles that are lethal in a homozygous individual may be carried in a heterozygote and thus maintained in a gene pool.

Two types of allele pairings can occur in the genotype: *homozygous* (pairing two of the same alleles, whether dominant, codominant, or recessive) and *heterozygous* (pairing of two different alleles). Recessive lethal alleles (e.g., Tay-Sachs disease) will, by definition, cause the death of only the homozygous recessive individual. Healthy heterozygous individuals will also contribute the masked recessive gene to the population's gene pool, allowing the gene to persist.

7. c. *Students know* new mutations are constantly being generated in a gene pool.

Mutation is an important source of genetic variation within a gene pool. These random changes take the form of additions, deletions, and substitutions of nucleotides and of rearrangements of chromosomes. The effect of many mutations is minor and neutral, being neither favorable nor unfavorable to survival and reproduction. Other mutations may be beneficial or harmful. The important principle is that culling, or selective breeding, cannot eliminate genetic diseases or unwanted traits from a population. The trait constantly reappears in the population in the form of new, spontaneous mutations.

7. d. *Students know* variation within a species increases the likelihood that at least some members of a species will survive under changed environmental conditions.

As environmental factors change, natural selection of adaptive traits must also be realigned. Variation within a species stemming either from mutation or from genetic recombination broadens the opportunity for that species to adapt to change, increasing the probability that at least some members of the species will be suitably adapted to the new conditions. Genetic diversity promotes survival of a species should the environment change significantly, and sameness can mean vulnerability that could lead to extinction.

7. e.* *Students know* the conditions for Hardy-Weinberg equilibrium in a population and why these conditions are not likely to appear in nature.

The principle of Hardy-Weinberg equilibrium, derived in 1908, asserts that the genetic structure of a nonevolving population remains constant over the generations. If mating in a large population occurs randomly without the influence of natural selection, the migration of genes from neighboring populations, or the occurrence of mutations, the frequency of alleles and of genotypes will remain constant over time. Such conditions are so restrictive that they are not likely to occur in nature precisely as predicted, but the Hardy-Weinberg equilibrium equation often gives an excellent approximation for a limited number of generations in sizeable, randomly mating populations. Even though genetic recombination is taken into account, mutations, gene flow between populations, and environmental changes influencing pressures of selection on a population do not cease to occur in the natural world.

7. f.* *Students know* how to solve the Hardy-Weinberg equation to predict the frequency of genotypes in a population, given the frequency of phenotypes.

The Hardy-Weinberg equilibrium equation can be used to calculate the frequency of alleles and genotypes in a population's gene pool. When only two alleles for a trait occur in a population, the letter p is used to represent the frequency of one allele, and the letter q is used to represent the frequency of the other. Students should agree first that the sum of the frequencies of the two alleles is 1, and this equation is written $p + q = 1$. That is, the combined frequencies of the alleles account for all the genes for a given trait.

Students should then consider the possible combinations of alleles in a *diploid organism* (the genome of a diploid organism consists of two copies of each chromosome). An individual could be homozygous for one allele (pp) or homozygous for the other (qq) or heterozygous (either pq or qp). These diploid genotypes will appear at frequencies that are the product of the allele frequencies (e.g., the frequency of a diploid pp individual is p^2, and the frequency of a diploid qq individual is q^2).

The heterozygotes are of two varieties, pq and qp (because the p allele might have been inherited from either parent), but the products of frequency pq and qp are the same. Therefore, the frequency of heterozygotes can simply be expressed as $2pq$. The sum of the frequencies of the homozygous and heterozygous individuals must equal 1, since all individuals have been accounted for. These principles are usually expressed as the equation $p^2 + 2pq + q^2 = 1$. Both equations represent different statements. The first ($p + q = 1$) is an accounting of the two types of alleles in the population, and the second ($p^2 + 2pq + q^2 = 1$) is an accounting of the three distinguishable genotypes.

If the allele frequencies are known (e.g., if $p = 0.1$ and $q = 0.9$) and Hardy-Weinberg equilibrium is assumed, then the frequencies p^2, $2pq$, and q^2 are

respectively 0.01, 0.18, and 0.81. That is, 81 percent of individuals would be homozygous *qq*. If *p* were a dominant (but nonselective) allele, then $p^2 + 2pq$, or 19 percent of the population, would express the dominant phenotype of the *p* allele.

The calculation can be used in reverse as well. If Hardy-Weinberg equilibrium conditions exist and 81 percent of the population expresses the *qq* recessive phenotype, then the allele frequency *q* is the square root of 0.81, and the rest of the terms can be calculated in a straightforward fashion.

Students can convince themselves of the state of equilibrium by constructing a Punnett Square that assumes random mating. The scenario might be a mass spawning of fish, in which 100,000 eggs and sperm are mixed in a stream and meet with each other randomly to form zygotes. Students can calculate the fraction of *p* and *q* type gametes in the stream by thinking through the types of gametes produced by heterozygous and homozygous adult fish. (For this exercise to work, the genotype distribution of adults must agree with Hardy-Weinberg equilibrium.) With the frequencies or numbers of each type of zygote calculated in the cells of a Punnett Square, students will see that equilibrium is preserved. Frequencies of alleles and genotypes, which are the genetic structure of the study population, would remain constant for generations under the premise of Hardy-Weinberg equilibrium.

STANDARD SET 8. Evolution (Speciation)

See the discussion introducing the concept of biological evolution that appears under Standard Set 7, "Evolution (Population Genetics)," on page 237.

8. Evolution is the result of genetic changes that occur in constantly changing environments. As a basis for understanding this concept:

a. *Students know* how natural selection determines the differential survival of groups of organisms.

Genetic changes can result from gene recombination during gamete formation and from mutations. These events are responsible for variety and diversity within each species. Natural selection favors the organisms that are better suited to survive in a given environment. Those not well suited to the environment may die before they can pass on their traits to the next generation. As the environment changes, selection for adaptive traits is realigned with the change. Traits that were once adaptive may become disadvantageous because of change.

Students can explore the process of natural selection further with an activity based on predator-prey relationships. The main purpose of these activities is to simulate survival in predator or prey species as they struggle to find food or to escape being consumed themselves. The traits of predator and prey individuals can be varied to test their selective fitness in different environmental settings.

An example of natural selection is the effect of industrial "melanism," or darkness of pigmentation, on the peppered moths of Manchester, England. These moths come in two varieties, one darker than the other. Before the industrial revolution, the dark moth was rare; however, during the industrial revolution the light moth seldom appeared. Throughout the industrial revolution, much coal was burned in the region, emitting soot and sulfur dioxide. For reasons not completely understood, the light-colored moth had successfully adapted to the cleaner air conditions that existed in preindustrial times and that exist in the region today.

However, the light-colored moth appears to have lost its survival advantage during times of heavy industrial air pollution. One early explanation is that when soot covered tree bark, light moths became highly visible to predatory birds. Once this change happened, the dark-peppered moth had an inherited survival advantage because it was harder to see against the sooty background. This explanation may not have been the cause, and an alternative one is that the white-peppered moth was more susceptible to the sulfur dioxide emissions of the industrial revolution. In any case, in the evolution of the moth, mutations of the genes produced light and dark moths. Through natural selection the light moth had an adaptive advantage until environmental conditions changed, increasing the population of the dark moths and depleting that of the light moths.

8. b. *Students know* a great diversity of species increases the chance that at least some organisms survive major changes in the environment.

This standard is similar to the previous standard set on diversity within a species but takes student understanding one step further by addressing diversity among and between species. For the same reasons pertinent to those for intraspecies diversity, increased diversity among species increases the chances that some species will adapt to survive future environmental changes.

8. c. *Students know* the effects of genetic drift on the diversity of organisms in a population.

If a small random sample of individuals is separated from a larger population, the gene frequencies in the sample may differ significantly from those in the population as a whole. The shifts in frequency depend only on which individuals fall in the sample (and so are themselves random). Because a random shift in gene frequency is not guaranteed to make the next generation better adapted, the shift—or genetic drift—with respect to the original gene pool is not necessarily an adaptive change. The *bottleneck effect* (i.e., nonselective population reductions due to disasters) and the *founder effect* (i.e., the colonization of a new habitat by a few individuals) describe situations that can lead to genetic drift of small populations.

8. d. *Students know* reproductive or geographic isolation affects speciation.

Events that lead to reproductive isolation of populations of the same species cause new species to appear. Barriers to reproduction that prevent mating between

populations are called *prezygotic* (before fertilization) if they involve such factors as the isolation of habitats, a difference in breeding season or mating behavior, or an incompatibility of genitalia or gametes. *Postzygotic* (after fertilization) barriers that prevent the development of viable, fertile hybrids exist because of genetic incompatibility between the populations, hybrid sterility, and hybrid breakdown.

These isolation events can occur within the geographic range of a parent population *(sympatric speciation)* or through the geographic isolation of a small population from its parent population *(allopatric speciation).* Sympatric speciation is much more common in plants than in animals. Extra sets of chromosomes, or *polypoidy,* that result from mistakes in cell division produce plants still capable of long-term reproduction but animals that are incapable of that process because polypoidy interferes with sex determination and because animals, unlike most plants, are usually of one sex or the other. Allopatric speciation occurs in animal evolution when geographically isolated populations adapt to different environmental conditions. In addition, the rate of allopatric speciation is faster in small populations than in large ones because of greater genetic drift.

8. e. *Students know* how to analyze fossil evidence with regard to biological diversity, episodic speciation, and mass extinction.

Analysis of the fossil record reveals the story of major events in the history of life on earth, sometimes called *macroevolution,* as opposed to the small changes in genes and chromosomes that occur within a single population, or *microevolution.* Explosive radiations of life following mass extinctions are marked by the four eras in the geologic time scale: the Precambrian, Paleozoic, Mesozoic, and Cenozoic. The study of biological diversity from the fossil record is generally limited to the study of the differences among species instead of the differences within a species. Biological diversity within a species is difficult to study because preserved organic material is rare as a source of DNA in fossils.

Episodes of speciation are the most dramatic after the appearance of novel characteristics, such as feathers and wings, or in the aftermath of a mass extinction that has cleared the way for new species to inhabit recently vacated adaptive zones. Extinction is inevitable in a changing world, but examples of mass extinction from the fossil record coincide with rapid global environmental changes. During the formation of the supercontinent Pangaea during the Permian period, most marine invertebrate species disappeared with the loss of their coastal habitats. During the Cretaceous period a climatic shift to cooler temperatures because of diminished solar energy coincided with the extinction of dinosaurs.

8. f.* *Students know* how to use comparative embryology, DNA or protein sequence comparisons, and other independent sources of data to create a branching diagram (cladogram) that shows probable evolutionary relationships.

The area of study that connects biological diversity to *phylogeny,* or the evolutionary history of a species, is called *systematics.* Systematic classification is based on

the degree of similarity between species. Thus, comparisons of embryology, anatomy, proteins, and DNA are used to establish the extent of similarities. Embryological studies reveal that *ontogeny,* development of the embryo, provides clues to phylogeny. In contrast to the old assertion that "ontogeny recapitulates phylogeny" (i.e., that it replays the entire evolutionary history of a species), new findings indicate that structures, such as gill pouches, that appear during embryonic development but are less obvious in many adult life forms may establish *homologies* between species (similarities attributable to a common origin). These homologies are evidence of common ancestry. Likewise, homologous anatomical structures, such as the forelimbs of humans, cats, whales, and bats, are also evidence of a common ancestor. Similarity between species can be evaluated at the molecular level by comparing the amino acid sequences of proteins or the nucleotide sequences of DNA strands. DNA-DNA hybridization, restriction mapping, and DNA sequencing are powerful new tools in systematics.

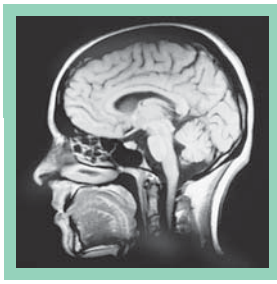
Approaches for using comparison information to classify organisms on the basis of evolutionary history differ greatly. *Cladistics* uses a branching pattern, or *cladogram,* based on shared derived characteristics to map the sequence of evolutionary change. The cladogram is a dichotomous tree that branches to separate those species that share a derived characteristic, such as hair or fur, from those species that lack the characteristic. Each new branch of the cladogram helps to establish a sequence of evolutionary history; however, the extent of divergence between species is unclear from the sequence alone.

Phenetics classifies species entirely on the basis of measurable similarities and differences with no attempt to sort homology from analogy. In recent years phenetic studies have been helped by the use of computer programs to compare species automatically across large numbers of traits. Striking a balance between these two approaches to classification has often involved subjective judgments in the final decision of taxonomic placement. Students can study examples of cladograms and create new ones to understand how a sequence of evolutionary change based on shared derived characteristics is developed.

8. g.* *Students know* how several independent molecular clocks, calibrated against each other and combined with evidence from the fossil record, can help to estimate how long ago various groups of organisms diverged evolutionarily from one other.

Molecular clocks are another tool to establish phylogenetic sequences and the relative dates of phylogenetic branching. Homologous proteins, such as cytochrome c, of different *taxa* (plants and animals classified according to their presumed natural relationships) and the genes that produce those proteins are assumed to evolve at relatively constant rates. On the basis of that assumption, the number of amino acid or nucleotide substitutions provides a record of change proportional to the time between evolutionary branches. The estimates of rate of change derived from these molecular clocks generally agree with parallel data from the fossil record; however, the branching orders and times between branches are more reliably determined by

measuring the degree of molecular change than by comparing qualitative features of morphology. When gaps in the fossil record exist, phylogenetic branching dates can be estimated by calibrating molecular change against the timeline determined from the fossil record.



STANDARD SET 9. Physiology (Homeostasis)

From the individual cell to the total organism, each functioning unit is organized according to *homeostasis,* or how the body and its parts deal with changing demands while maintaining a constant internal environment. In 1859 noted French physiologist Claude Bernard described the difference between the internal environment of the cells and the external environment in which the organism lives. Organisms are shielded from the variations of the external environment by the "constancy of the internal milieu." This "steady state" refers to the dynamic equilibrium achieved by the integrated functioning of all the parts of the organism.

American physiologist Walter Cannon called this phenomenon *homeostasis,* which means "standing still." All organ systems of the human body contribute to homeostasis so that blood and tissue constituents and values stay within a normal range. Students will need supportive review of the major systems of the body and of the organ components of those systems (see Standard Set 2, "Life Sciences," for grade five in Chapter 3 and Standard Set 5, "Structure and Function in Living Systems," for grade seven in Chapter 4). As the prime coordinators of the body's activities, the nervous and endocrine systems must be examined and their interactive roles clearly defined.

9. As a result of the coordinated structures and functions of organ systems, the internal environment of the human body remains relatively stable (homeostatic) despite changes in the outside environment. As a basis for understanding this concept:

a. *Students know* how the complementary activity of major body systems provides cells with oxygen and nutrients and removes toxic waste products such as carbon dioxide.

The digestive system delivers nutrients (e.g., glucose) to the circulatory system. Oxygen molecules move from the air to the alveoli of the lungs and then to the circulatory system. From the circulatory system glucose and oxygen molecules move from the capillaries into the cells of the body where cellular respiration occurs. During cellular respiration these molecules are oxidized into carbon dioxide and water, and energy is trapped in the form of ATP. The gas exchange process is reversed for the removal of carbon dioxide from its higher concentration in the cells to the circulatory system and, finally, to its elimination by exhalation from the lungs.

The concentration of sugar in the blood is monitored, and students should know that sugar can be stored or pulled from reserves (glycogen) in the liver and

muscles to maintain a constant blood sugar level. Amino acids contained in proteins can also serve as an energy source, but first the amino acids must be deaminated, or chemically converted, in the liver, producing ammonia (a toxic product), which is converted to water-soluble urea and excreted by the kidneys. Teachers should emphasize that all these chemicals are transported by the circulatory system and the cells. Organs at the final destination direct these chemicals to their exit from the circulatory system.

9. b. *Students know* how the nervous system mediates communication between different parts of the body and the body's interactions with the environment.

An individual becomes aware of the environment through the sense organs and other body receptors (e.g., by allowing for touch, taste, and smell and by collecting information about temperature, light, and sound). The body reflexively responds to external stimuli through a reflex arc (see Standard 9.e in this section). (A *reflex arc* is the pathway along the central nervous system where an impulse must travel to bring about a reflex; e.g., sneezing or coughing.) Students can examine the sense organs, identify other body receptors that make them aware of their environment, and see ways in which the body reflexively responds to an external stimulus through a reflex arc.

Hormones work in conjunction with the nervous system, as shown, for example, in the digestive system, where insulin released from the pancreas into the blood regulates the uptake of glucose by muscle cells. The pituitary master gland produces growth hormone for controlling height. Other pituitary hormones have specialized roles (e.g., follicle-stimulating hormone [FSH] and luteinizing hormone [LH] control the gonads, thyroid-stimulating hormone [TSH] controls the thyroid, and adrenocorticotropic hormone [ACTH] regulates the formation of glucocorticoids by the adrenal cortex). This master gland is itself controlled by the hypothalamus of the brain.

9. c. *Students know* how feedback loops in the nervous and endocrine systems regulate conditions in the body.

Feedback loops are the means through which the nervous system uses the endocrine system to regulate body conditions. The presence or absence of hormones in blood brought to the brain by the circulatory system will trigger an attempt to regulate conditions in the body. To make feedback loops relevant to students, teachers can discuss the hormone leptin, which fat cells produce as they become filled with storage reserves. Leptin is carried by the blood to the brain, where it normally acts to inhibit the appetite center, an example of negative feedback. When fat reserves diminish, the concentration of leptin decreases, a phenomenon that in turn causes the appetite center in the brain to start the hunger stimulus and activate the urge to eat.

9. d. *Students know* the functions of the nervous system and the role of neurons in transmitting electrochemical impulses.

Transmission of nerve impulses involves an electrochemical "action potential" generated by gated ion channels in the membrane that make use of the countervailing gradients of sodium and potassium ions across the membrane. Potassium ion concentration is high inside cells and low outside; sodium ion concentration is the opposite. The sodium and potassium ion concentration gradients are restored by an active transport system, a pump that exchanges sodium and potassium ions across the membrane and uses ATP hydrolysis as a source of free energy. The release of neurotransmitter chemicals from the axon terminal at the synapse may initiate an action potential in an adjacent neuron, propagating the impulse to a new cell.

9. e. *Students know* the roles of sensory neurons, interneurons, and motor neurons in sensation, thought, and response.

The pathways of impulses from dendrite to cell body to axon of sensory neurons, interneurons, and motor neurons link the chains of events that occur in a reflex action. Students should be able to diagram this pathway. Similar paths of neural connections lead to the brain, where the sensations become conscious and conscious actions are initiated in response to external stimuli. Students might also trace the path of the neural connections as the sensation becomes conscious and a response to the external stimulus is initiated. Students should also be able to identify gray and white matter in the central nervous system.

9. f.* *Students know* the individual functions and sites of secretion of digestive enzymes (amylases, proteases, nucleases, lipases), stomach acid, and bile salts.

To bring about digestion, secretions of enzymes are mixed with food (in the mouth and as the food proceeds from the mouth through the stomach and through the small intestines). For example, salivary glands and the pancreas secrete amylase enzymes that change starch into sugar. Stomach acid and gastric enzymes begin the breakdown of protein, a process that intestinal and pancreatic secretions continue.

Lipase enzymes secreted by the pancreas break down fat molecules (which contain three fatty acids) to free fatty acids plus diglycerides (which contain two fatty acids) and monoglycerides (which contain one fatty acid). Bile secreted by the liver furthers the process of digestion, emulsifying fats and facilitating digestion of lipids. Students might diagram the digestive tract, labeling important points of secretion and tracing the pathways from digestion of starches, proteins, and other foods. They can then outline the role of the kidney nephron in the formation of urine and the role of the liver in glucogenesis and glycogenolysis (glucose balance) and in blood detoxification.

9. g.* *Students know* the homeostatic role of the kidneys in the removal of nitrogenous wastes and the role of the liver in blood detoxification and glucose balance.

Microscopic nephrons within the kidney filter out body wastes, regulate water, and stabilize electrolyte levels in blood. The liver removes toxic materials from the blood, stores them, and excretes them into the bile. The liver also regulates blood glucose.

9. h.* *Students know* the cellular and molecular basis of muscle contraction, including the roles of actin, myosin, Ca^{+2}, and ATP.

Controlled by calcium ions and powered by hydrolysis of ATP, actin and myosin filaments in a sarcomere generate movement in stomach muscles. Striated muscle fibers reflect the filamentous makeup and contraction state evidenced by the banding patterns of those fibers. A sketch of the sarcomere can be used to indicate the functions of the actin and myosin filaments and the role of calcium ions and ATP in muscle contraction.

9. i.* *Students know* how hormones (including digestive, reproductive, osmoregulatory) provide internal feedback mechanisms for homeostasis at the cellular level and in whole organisms.

Hormones act as chemical messengers, affecting the activity of neighboring cells or other target organs. Their movement can be traced from their point of origin to the target site. The feedback mechanism works to regulate the activity of hormones and promotes homeostasis.

STANDARD SET 10. Physiology (Infection and Immunity)

Some bacteria, parasites, and viruses cause human diseases because they either rob the body of necessary sustenance or secrete toxins that cause injury. The human body has a variety of mechanisms to interfere with or destroy invading pathogens. Besides protection afforded by the skin, one of the most effective means of defending against agents that harm the body is the immune system with its cellular and chemical defenses. Students should develop a clear understanding of the components of the immune system and know how vaccines and antibiotics are used to combat disease. They should also know that acquired immune deficiency syndrome (AIDS) compromises the immune system, causing affected persons to succumb to other AIDS-associated infections that are harmless to people with an intact immune system.

10. Organisms have a variety of mechanisms to combat disease. As a basis for understanding the human immune response:

a. *Students know* the role of the skin in providing nonspecific defenses against infection.

The skin serves as a physical barrier to prevent the passage of many disease-causing microorganisms. Cuts and abrasions compromise the skin's ability to act as a barrier. Teachers can use charts and overhead projections to show the dangers and physiologic responses of a break in the skin.

10. b. *Students know* the role of antibodies in the body's response to infection.

Cells produce antibodies to oppose antigens, substances that are foreign to the body. An example of an antigen is a surface protein of a flu virus, a protein with a shape and structure unlike those of any human proteins. The immune system recognizes that the flu virus structure is different and generates proteins called *antibodies* that bind to the flu virus. Antibodies can inactivate pathogens directly or signal immune cells that pathogens are present.

10. c. *Students know* how vaccination protects an individual from infectious diseases.

Several weeks are required before the immune system develops immunity to a new antigen. To overcome this problem, vaccinations safely give the body a look in advance at the foreign structures. Vaccines usually contain either weakened or killed pathogens that are responsible for a specific infectious disease, or they may contain a purified protein or subunit from the pathogen. Although the vaccine does not cause an infectious disease, the antigens in the mixture prompt the body to generate antibodies to oppose the pathogen. When the individual is exposed to the pathogenic agent, perhaps years later, the body still remembers having seen the antigens in the vaccine dose and can respond quickly. Students have been exposed to the practical aspects of immunization through their knowledge of the vaccinations they must receive before they can enter school. They have all experienced getting shots and may have seen their personal vaccination record in which dates and kinds of inoculations are recorded. The review of a typical vaccination record, focusing on the reason for the shots and ways in which they work, may serve as an effective entry to the subject.

Students should review the history of vaccine use. Early literature provides descriptions of vaccine use from pragmatic exposure, but the term *vaccine* is derived from the cowpox exudate that Edward Jenner used during the 1700s to inoculate villagers against the more pathogenic smallpox. Louis Pasteur, noted for his discovery of the rabies treatment, also developed several vaccines. Poliovirus, the cause of infantile paralysis (poliomyelitis), was finally conquered in the 1950s through vaccines that Jonas Salk and Albert B. Sabin refined.

10. d. *Students know* there are important differences between bacteria and viruses with respect to their requirements for growth and replication, the body's primary defenses against bacterial and viral infections, and effective treatments of these infections.

A *virus,* which is the simplest form of a genetic entity, is incapable of metabolic life and reproduction outside the cells of other living organisms. A virus contains genetic material but has no ribosomes. Although some viruses are benign, many harm their host organism by destroying or altering its cell structures. Generally, the body perceives viruses as antigens and produces antibodies to counteract the virus. *Bacteria* are organisms with a full cellular structure. They, too, can be benign or harmful. Harmful bacteria and their toxins are perceived as antigens by the body, which in turn produces antibodies. In some cases infectious diseases may be treated effectively with *antiseptics,* which are chemicals that oxidize or in other ways inactivate the infecting organism. Antiseptics are also useful in decontaminating surfaces with which the body may come in contact (e.g., countertops). *Antibiotics* are effective in treating bacterial infections, sometimes working by destroying or interfering with the growth of bacterial cell walls or the functioning of cell wall physiology or by inhibiting bacterial synthesis of DNA, RNA, or proteins. Antibiotics are ineffective in treating viral infections.

Students might research infections caused by protists (malaria, amoebic dysentery), bacteria (blood poisoning, botulism, food poisoning, tuberculosis), and viruses (rabies, colds, influenza, AIDS). They might also investigate the pathogens currently being discussed in the media and study each infectious organism's requirements for growth and reproduction. Teachers should review the dangers of common bacteria becoming resistant to antibiotics through long-standing overapplication, as shown by the increasing incidence of drug-resistant tuberculosis and other bacteria. Using a commercially available kit, teachers can demonstrate how antibiotics may act generally or specifically against bacteria. Agar plates may be inoculated with different bacteria, and different antibiotic discs may be placed on these plates to create a clear zone in which growth around the antibiotic discs is inhibited.

10. e. *Students know* why an individual with a compromised immune system (for example, a person with AIDS) may be unable to fight off and survive infections by microorganisms that are usually benign.

When an immune system is compromised (e.g., through infection by the human immunodeficiency virus [HIV]), it becomes either unable to recognize a dangerous antigen or incapable of mounting an appropriate defense. This situation happens when the virus infects and destroys key cells in the immune system.

10. f.* *Students know* the roles of phagocytes, B-lymphocytes, and T-lymphocytes in the immune system.

Phagocytes move, amoebalike, through the circulatory system, consuming waste and foreign material, such as aged or damaged blood cells and some infectious bacteria and viruses. Two broad types of lymphocytes (a class of white blood cells) originate in the bone marrow during embryonic life. One type (the B-lymphocyte) matures in the bone marrow and gives rise to antibody-producing plasma cells that are responsible for humoral immunity. Each mature B-lymphocyte can give rise to only a single type of antibody, which itself may recognize only a single foreign antigen. The other type (the T-lymphocyte) matures in the thymus gland during embryogenesis and gives rise to "cytotoxic" (cell killing) and "helper" T-lymphocytes. The cytotoxic T-cells are particularly useful for surveillance of intra-cellular pathogens. Antibodies cannot reach the intracellular pathogen because of the cell membrane, but the infected cell can be identified and killed. Helper T-cells assist in organizing both the humoral and cellular immune responses.